Unit 1: Fundamentals of Chemistry
Exercise Questions

Exercise Multiple Choice Question Answers

1. **Industrial chemistry deals with the manufacturing of compounds:**
   (a) in the laboratory (b) on micro scale
   (c) on commercial scale (d) on economic scale
2. **Which one of the following can be separated by physical means?**
   (a) mixture (b) element (c) compound (d) radical
3. **The most abundant element occurring in the oceans is:**
   (a) oxygen (b) hydrogen (c) nitrogen (d) silicon
4. **Which one of the following element is found in most abundance in the earth's crust?**
   (a) oxygen (b) aluminum (c) silicon (d) iron
5. **The third abundant gas found in the atmosphere is:**
   (a) carbon monoxide (b) oxygen. (c) nitrogen (d) argon
6. **One amu (atomic mass unit) is equivalent to:**
   (a) 1.66×10^{-24} mg (b) 1.66×10^{-24}g (c). 1.66×10^{-21}g (d) 1.66×10^{-23}g
7. **All of the followings are tri-atomic molecule except:**
   (a) H_2 (b) O_3 (c) H_2O (d) CO_2
8. **The mass of one molecule of water is:**
   (a) 18 amu (b) 18 g (c) 18 mg (d) 18 kg
9. **The molar mass of H_2S0_4 is·**
   (a) 98 g (b) 48 amu (c) 4.8 g (d) 98 amu
10. **Molar mass is usually expressed in grams. Which one of the following is molar mass of O_2 in amu?**
    (a) 32 amu (b) 53.2×10^{-24} amu
    (c) 1.92×10^{-25} amu (d) 192.64×10^{-25} amu
11. **How many numbers of moles are equivalent to 8 grams of CO_2?**
    (a) 0.15 (b) 0.18 (c) 0.21 (d) 0.24
12. **Which one of the following pairs has the same number of ions?**
    (a) 1 mole of NaCl and 1 mole of $MgCl_2$
    (b) 1/2 mole of NaCl and 1/2 mole of $MgCl_2$
    (c) 1/2 mole of NaCl and 1/3 mole of $MgCl_2$
    (d) 1/3 mole of NaCl and 1/2 mole of $MgCl_2$
13. **Which one of the following pairs has the same mass?**
    (a) 1 mole of CO and 1 mole of N_2 (b) 1 mole of CO and 1 mole of CO_2
    (c) 1 mole of O_2 and 1 mole of N_2 (d) 1 mole of O_2 and 1 mole of CO_2

ANSWR KEY

1	c	3	a	5	d	7	a	9	a	11	b	13	a
2	a	4	a	6	b	8	a	10	a	12	c		

Exercise Short Questions Answers

Q.1 Define industrial chemistry and analytical chemistry.

Ans: **Industrial Chemistry**

"The branch of chemistry that deals with the manufacturing of chemical substances (elements and compounds) on commercial scale, is called industrial chemistry."

Applications:

i. It deals with the manufacturing of basic chemicals such as oxygen, chlorine, ammonia, caustic soda, nitric acid and sulphuric acid.

ii. Use of these chemicals to provide the raw materials for many other industries such as fertilizers, soap, textiles, agricultural products, paints and paper etc

Analytical Chemistry

"The branch of chemistry that deals with separation and analysis of a sample to identify its components is called analytical chemistry. The separation is carried out prior to qualitative and quantitative analysis."

Qualitative Analysis

It provides the identity of a substance (composition of chemical species).

Quantitative Analysis

It determines the amount of each component present in the sample.

Scope

In this branch different techniques and instruments used for analysis are studied.

The scope of this branch covers food, water, environmental and clinical analyses.

Q.2 How can you differentiate between organic and inorganic chemistry?

Ans: **Organic Chemistry**

"The branch of chemistry that deals with the study of covalent compounds of carbon and hydrogen and their derivatives is called organic chemistry."

Scope

Organic chemists determine the structure and properties of these naturally occurring as well as synthesized compounds.

Scope of this branch covers petroleum, petrochemicals and pharmaceutical industries.

Inorganic Chemistry

"The branch of chemistry that deals with the study of all elements and their compounds except those of compounds of carbon and hydrogen (hydrocarbons) and their derivatives is called inorganic chemistry."

Applications

It has applications in every aspect of the chemical industry such as glass, cement, ceramics and metallurgy (extraction of metals from ores).

Q.3 Give the scope of biochemistry.

Ans: *"The branch of chemistry that deals with the study of structure, composition, and chemical reactions of substances found in living organism."*

Scope

It covers all chemical processes taking place in living organisms such as synthesis and metabolism of bio-molecules like carbohydrates, proteins, fats etc.

Emergence of biochemistry as a separate discipline

Biochemistry emerged as a separate discipline when scientists began to study:

i. How living things obtain energy from food?

ii. How' the fundamental biological changes occur during a disease?

Applications:

It is applied in the fields of medicine, food science and agriculture.

Q.4 Wow does homogeneous mixture differ from heterogeneous mixture?

Ans:

Homogeneous Mixture	Heterogeneous Mixture
Mixtures that have uniform composition throughout are called homogeneous mixtures. *It is called solution.*	*Those mixtures in which composition is not uniform throughout are called heterogeneous mixtures.*
For example: Air, gasoline and ice cream	**For example:** Soil, rock, wood, concrete and paint etc.

Q.5 What is the relative atomic mass? How it is related to gram?

Ans: **Relative atomic mass**

"The average mass of an atom of an element as compared to $1/12^{th}$ (one-twelfth) the mass of one atom of carbon-12 isotope is called relative atomic mass."

Unit of relative atomic mass:

Its unit is atomic mass unit, with symbol amu.

Atomic mass unit:

"One atomic mass unit is $1/12^{th}$ the mass of one atom of carbon-12^{th}."

When this atomic mass unit is expressed in grams it is.

$$1 \text{ amu} = 1.66 \times 10^{-24} \text{ g}$$

Q.6 Define empirical formula with example.

Ans: **Empirical Formula:**

"It is the simplest whole number ratio of atoms present in a compound."

The empirical formula of a compound is determined by knowing the percentage composition of a compound.

Example:

Glucose has simplest ratio 1: 2: 1 of carbon, hydrogen and oxygen respectively.

Hence its empirical formula is CH_2O.

Q.7 State three reason why do you think air is a mixture and water is a compound?

Ans: **Reasons:**

i. Water is a compound because it is formed by chemical combination of hydrogen and oxygen whereas air is formed by simple mixing of different gases.

ii. Water has fixed ratio between masses of hydrogen and oxygen, whereas in air ratio between masses of component gases is not fixed.

iii. Water has definite melting and boiling points whereas air does not have any fixed melting and boiling point.

Q.8 Explain why are hydrogen and oxygen considered elements whereas water as a compound.

Ans: Hydrogen and oxygen are elements because they have same type of atoms, having same atomic number and it cannot be decompose into simple substances by chemical means. Water is considered as compound because it is a substance made up of two or more elements chemically combined together in a fixed ratio by mass. As a result of this combination oxygen and hydrogen lose their own properties and produce new substance (H_2O).

Q.9 What is the significance of the symbol of an element?

Ans: **Significance of the symbol of an element:**

Symbols are used for elements instead of writing of their complete names. So, it takes less time/save time and element can be recognized by that symbol in all over the world.

i. Symbol represents the name of an element.

ii. It represents one atom of the element

iii. It represents one mole of atoms of the element.

iv. It represents atomic mass of an element.

v. It helps to write and understand chemical equation for different chemical reactions.

vi. Periodic table is based on symbols of different elements.

For example: Oxygen (O), Sulphur (S), Nitrogen (N)

Q.10 State the reasons: soft drink is a mixture and water is a compound.

Ans:

Mixture (Soft Drink)	Compound (Water)
• Soft drink is made up of simple mixing up of substances without any fixed ratio.	• Water is formed by chemical combination of atoms of elements hydrogen and oxygen in a fixed ratio of 1:8 by mass.
• Soft drink has heterogeneous composition.	• Water has homogeneous composition.
• Its components can be separated by physical means.	• Its components can't be separated by physical means

Q.11 Classify the following into element, compound and mixture:

- **He and H_2**
- **CO and CO_2**
- **Water and milk**
- **Gold and brass**
- **Iron and steel**

Ans:

(i) He and H_2: He and H_2 are elements

(ii) CO and Co: CO is a compound and Co is an element.

(iii) Water and milk: Water is a compound and milk is a mixture

(iv) Gold and brass: Gold is an element and brass is a mixture

(v) Iron and steel: Iron is an element and steel is a mixture.

Q.12 Define atomic mass unit. Why is it needed?

Ans: **<u>Atomic mass unit</u>**

"The mass equal to one twelfth of the mass of a carbon -12 atom is called atomic mass unit."

The atomic mass unit is abbreviated as amu.

1 amu = 1/12 × mass of C-12 atom

The mass of one atom of carbon -12 is 12 amu.

It is the unit used for the relative atomic mass. It is used to compare masses of atoms.

Q.13 State the nature and name of the substance formed by combining the following:
i. Zinc + Copper ii. Water + Sugar iii. Aluminium + Sulphur
iv. Iron + Chromium + Nickel

Ans:

(i) Zinc + Copper

It is a mixture or alloy. The name of alloy is brass.

(ii) Water + Sugar

It is a mixture. The name of mixture or solution is syrup.

(iii) Aluminium + Sulphur

It forms compound. The name of compound is aluminium -sulphide.

(iv) Iron + Chromium + Nickel

It is a mixture or alloy. The name of alloy is nichrome.

Q.14 Differentiate between molecular mass and formula mass, which of the following will be molecular formula?

- H_2O
- NaCl
- KI
- H_2SO_4

Ans:

Molecular Mass	Formula Mass
i. The sum of atomic masses of all the atoms present in one molecule of a molecular substance called molecular mass. **ii.** The term molecular mass is used for compounds that exist as molecules.	**i.** *The sum of atomic masses of all the atoms present in one formula unit of an ionic compound is called formula mass.* **ii.** The term formula mass is used for compounds that exits as formula units i.e. the compounds consists of ions.
For example: Molecular mass of water is 18 amu and that of carbon is 44 amu	**For example:** Formula mass of sodium chloride is 58.5 amu and that of $CaCO_3$ is 100 amu.

H_2O and H_2SO_4 are the molecular formula.

Q.15 Which has more then atoms: 10 g of *Al* or 10 g of Fe?

Ans: 10 g of Al has more atoms than 10 g of Fe.

For Al

i. Given mass of Al = 10g

Molar mass of Al = 27 g mol^{-1}

No of atoms in 10g of Al = No of moles × N_A

$$\text{Number of atom} = \frac{\text{Mass}}{\text{Molar Mass}} \times N_A$$

$$= \frac{10}{23} \times 6.02 \times 10^{23}$$

$$= 2.23 \times 10^{23} \text{ atoms}$$

ii. Given mass of Fe = 10g

Molar mass of Fe = 56 g mol^{-1}

For Fe

$$\text{Number of atom} = \frac{\text{Mass}}{\text{Molar Mass}} \times N_A$$

$$= \frac{10}{56} \times 6.02 \times 10^{23}$$

$$= 1.075 \times 10^{23}$$

Therefore: Aluminium has more number of atoms than iron.

Q.16 Which one has more molecules: 9 g of water or 9 g of sugar ($C_{12}H_{22}O_{11}$)?

Ans: 9 g of H_2O has more molecules than 9 g of $C_6H_{12}O_6$

i. Given mass of water (H_2O) = 9 g

Molar mass of water (H_2O) = 18 g mol^{-1}

Number of molecules in 9g of water $= \frac{\text{Given Mass}}{\text{Molar Mass}} \times N_A$

$$= \frac{9}{18} \times 6.02 \times 10^{23} \text{ molecules}$$

$$= 3.01 \times 10^{23} \text{ molecules}$$

ii. Given mass of sugar = 9 g

Molar mass of sugar = 342 g mol^{-1}

Number of molecules in 9g of sugar $= \frac{\text{Mass}}{\text{Molar Mass}} \times N_A$

$$= \frac{9}{342} \times 6.02 \times 10^{23}$$

$$= 1.584 \times 10^{22} \text{ molecules}$$

Therefore: 9 g of H_2O has more molecules than 9 g of $C_{12}H_{22}O_1$.

Q.17 Which one has more formula units: 1 g of Na*Cl* or 1 g of K*Cl*?

Ans:

i. Given mass = 1g

Formula mass of NaCl = 58.5 g mol^{-1}

Formula units in 1g of NaCl $= \frac{\text{Given mass}}{\text{Formula mass}} \times N_A$

$$= \frac{1}{58.5} \times 6.02 \times 10^{23}$$

$$= 1 - 029 \times 10^{22} \text{ formula units}$$

ii. Given mass of KCl = 1g

Formula mass of KCl = 74.5 g mol^{-1}

Formula units $= \frac{\text{Given mass}}{\text{Formula mass}} \times N_A$

$$= \frac{1}{74.5} \times 6.02 \times 10^{23}$$

$$= 8.080 \times 10^{21} \text{ formula units}$$

Therefore 1g of NaC*l* has more formula units than 1g of KC*l*.

Q.18 Differentiate between homoatomic and heteroatomic molecules with examples.

Ans:

Homoatomic molecules	Heteroatomic molecules
A molecule containing same type of atoms is called homoatomic molecule.	*A molecule consists of different kinds of atoms is called a heteroatomic molecule.*
Examples: Hydrogen (H_2), Oxygen (O_3) and sulphur (S_8) These are molecules of elements	**Examples:** CO_2, H_2O and NH_3 These are molecules of compounds

Q.19: In which one of the followings the number of hydrogen atoms in more? 2 moles of HCl or 1 mole of NH_3 (Hint: 1 mole of a substance contains as much number of moles of atoms as are in 1 molecule of a substance.)

Ans: No of moles of hydrogen in 1 mole of HCl = 1 mole

No of moles of hydrogen in 2 moles of HCl = 2 moles

Whereas No of moles of hydrogen in 1 mole of NH_3= 3 moles

Hence 1 mole of NH_3 contains 3 moles of hydrogen will have more hydrogen atoms than 2 moles of hydrogen present in 2 moles of HCl.

Exercise Long Question Answers

Q.1 Define element and classify the elements with examples.

Ans: See Q. No. **4** (Subjective Part, Long Questions Answers)

Q.2 List five characteristics by which compounds can be distinguished from mixtures.

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.3 Differentiate between the following with examples:

(a) Molecule and gram molecule **(b) Atom and gram atom**

(c) Molecular mass and molar mass **(d) Chemical formula and gram formula**

Ans: See Q. No. **21** (Subjective Part, Long Questions Answers)

Q.4 Mole is SI unit for the amount of a substance. Define it with examples?

Ans: See Q. No. **23** (Subjective Part, Long Questions Answers)

Exercise Solved Numericals

Q.1 Sulphuric acid is the king of chemicals. If you need 5 moles of sulphuric acid for a reaction, how many grams of it will you weigh?

Given Data:

Number of moles of H_2SO_4 = 5 moles

Molar mass of H_2SO_4 = $2(1)+1(32)+4(16)$

= 2 + 32 + 64

= 98 g/mol

To find:

Mass of H_2SO_4 = ?

Solution:

$$\text{Number of Moles} = \frac{\text{Mass of } H_2SO_4}{\text{Molar mass of } H_2SO_4}$$

Mass in grams = No of moles × molar mass = 5×98 = 490 g

Therefore, 5 moles of sulphuric acid will have mass 490 g.

Q.2 Calcium carbonate is insoluble in water. If you have 40 g of it; how m~ Ca^{2+} and CO_3^{2-} ions are present in it?

Given Data:

Mass of calcium carbonate = 40g

Molar mass of calcium carbonate = $CaCO_3$

= (40×1) + (12×1)+(16×3)

= 40 + 12 + 48

= 100 g/mol

To find:

Number of Ca^{2+} ions = ?

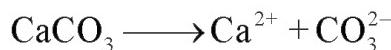
Number of CO_3^{2-} ions = ?

Solution:

$$\text{Number of Moles of } CaCO_3 = \frac{\text{Mass}}{\text{Molar mass}}$$

$$= \frac{40}{100} = 0.4 \text{ mol}$$

Balanced equation:

$$CaCO_3 \longrightarrow Ca^{2+} + CO_3^{2-}$$

Number of moles of $CaCO_3$ = 0.4 mole

Number of moles of Ca^{2+} ions in one mole of $CaCO_3$ = 6.02×10^{23}

Number of Ca+2 ion in 0.4 moles of $CaCO_3$ = No of moles × N_A

= $0.4\times6.02\times10^{23}$

= 2.408×10^{23} ions

We know that

No of Ca^{2+} ions = No of CO_3^{2-} ions

No of CO_3^{2-} = 2.40×10^{23} ions

= 2.408 × 10^{23} ions

Q.3 If you have 6.02 × 10^{23} ions of aluminium; how many sulphate ions will be required to prepare $Al_2(SO_4)_3$?

Given Data:

Number of ions of Al^{3+} = 6.02×10^{23}

Number of sulphate ions in $Al_2(SO_4)_3$ =?

Solution:

$$Al^{3+} + 3SO_4^{2-} \longrightarrow Al_2(SO_4)_3$$

According to balanced chemical equation:

No of moles of SO_4^{2-} ion required for 2 moles of Al^{3+} ions = 3

No of moles of SO_4^{2-} ions for 1 mole of Al^{3+} = 3/2 = 1.5 moles

Thus, number of SO_4^{2-} ions = No of moles × N_A

= $1.5\times6.02\times10^{23}$

= 9.03×10^{23} ions

Q.4 Calculate the number of molecules of the following compounds:

a. 16 g of H_2CO_3 b. 20 g of HNO_3 c. 30 g of $C_6H_{12}O_6$

Given Data:

a. 16g of H_2CO_3

Given mass of H_2CO_3 = 16g

Molar mass of H_2CO_3 = 2(1)+1(12)+3(16)

= 2 + 12 + 48

= 62g / mol

Number of molecules of H_2CO_3 = ?

Number of molecules of H_2CO_3 $= \dfrac{\text{Givenmass of } H_2CO_3}{\text{Molar mass of } H_2CO_3} \times N_A$

$= \dfrac{16}{62} \times 6.02\times10^{23}$

= 1.505 × 10^{23} molecules

b. 20g of HNO_3

Given data:

Given mass of HNO_3 = 20g

Molar mass of HNO_3 = 1(1)+1(14)+3(16)

= 1+14+48=63g/mol

Number of molecules of HNO3 = ?

Solution:

Number of molecules of HNO_3 $= \frac{\text{Given mass of } HNO_3}{\text{Molar mass of } HNO_3} \times N_A$

$= \frac{20}{63} \times 6.02 \times 10^{23}$

$= 1.91 \times 10^{23}$ molecules

$= 1.908 \times 10^{23}$ molecules

c. 30g of $C_6 H_{12} O_6$

Given data:

Given mass of $C_6H_{12}O_6$ = 30g.

Molar mass of $C_6 H_{12} O_6$ = 72 + 12 + 96

= 180g / mol

Number of moles of $C_6H_{12}O_6$ $= \frac{\text{Given mass of } C_6H_{12}O_6}{\text{Molar mass of } C_6H_{12}O_6} \times N_A$

$= \frac{30}{180} \times 6.02 \times 10^{23}$

$= 1 \times 10^{23}$ molecules

Q.5 Calculate the number of ions in the following compounds:

a. 10g of $AlCl_3$ b. 30 g of $BaCl_2$ c. 58 g of H_2SO_4

Ans:

a. 10g of $AlCl_3$

Given Data:

Given mass of $AlCl_3$ = 10g

Molar mass of $AlCl_3$ = 1(27)+3(35.5)

= 133.5 g / mol

No. of ions of $AlCl_3$ = ?

No of formula units of $AlCl_3$ in 10g $= \frac{\text{Given mass of } AlCl_3}{\text{Molar mass of } AlCl_3}$

$= \frac{10}{133.5} \times 6.02 \times 10^{23}$

$= 0.451 \times 10^{23}$ formula units

1 formula unit of $AlCl_3$ contains total number of ions = 4ions

4.51×10^{22} formula units of $AlCl_3$ contain total number of ions $= 4 \times 0.451 \times 10^{23}$

$= 1.80 \times 10^{24}$ ions

Therefore number of ions in 10g of $AlCl_3$ = 1.80×10^{23} ions

b. 30g of $BaCl_2$

Given Data:

Given mass of $BaCl_2$ = 30g

Molar mass of $BaCl_2$ = 1(137)+2(35.5)

= 137+71=208 g/mol

No of ions of 30g of $BaCl_2$ = ?

$$\text{No. of formula units in 30g of } BaCl_2 = \frac{\text{Givenmass of } BaCl_2}{\text{Molar mass of } BaCl_2} \times N_A$$

$$= \frac{30}{208} \times 6.02\times10^{23} = 0.86\times10^{23} \text{ formula units}$$

1 formula unit of $BaCl_2$ contains total number of ions = 3

0.86×10^{23} formula units will contain total number of ions = $3\times0.86\times10^{23}$ ions

= 2.58×10^{23} ions

c. 58g of H_2SO_4

Given Data:

Given mass of H_2SO_4 = 58g

Molar mass of H_2SO_4 = 2(1)+1(32)+4(16)

=2+32+64

=98 g/mol

Number of ion of H_2SO_4 = ?

$$\text{Number of formula units in 58g of } H_2SO_4 = \frac{\text{Given mass of } H_2SO_4}{\text{Molar mass of } H_2SO_4} \times N_A$$

$$= \frac{58}{98} \times 6.02\times10^{23}$$

= 3.56×10^{23} formula units

1 formula unit of H_2SO_4 contains total number of ions = 3 ions

3.56×10^{23} formula units of H2SO4 contain total number of ions = $3\times3.56\times10^{23}$

= 10.682×10^{23} ions

Q.6 What will be the mass of 2.05×10^{16} molecules of H_2SO_4

Ans:

Given Data:

Number of molecules of H_2SO_4 = 2.05×10^{16}

Molar mass of H_2SO_4 = 2(1)+1(32)+4(18)=2+32+64=98g/mol

Mass of H_2SO_4 = ?

Solution:

$$\text{Number of molecules of } H_2SO_4 = \frac{\text{Mass of } H_2SO_4}{\text{Molar mass of } H_2SO_4} \times N_A$$

$$2.05\times10^{16} = \frac{\text{Mass of } H_2SO_4}{98} \times 6.02\times10^{23}$$

$$\text{Mass of } H_2SO_4 = \frac{2.05\times10^{16}\times98}{6.02\times10^{23}}$$

= 3.332×10^{-6}g

Q.7 How many total atoms are required to prepare 60 g of HNO_3?

Given Data:

Given mass of HNO_3 = 60g

Molar mass of HNO_3 = 1(1)+1(14)+3(16)

= 1 + 14 + 48

= 63g / mol

No of atoms of HNO_3 = ?

Solution:

$$\text{Number of molecules of } HNO_3 = \frac{\text{Given mass of } NHO_3}{\text{Molar mass of } HNO_3} \times N_A$$

$$= \frac{60}{63} \times 6.02 \times 10^{23}$$

$$= 0.95 \times 6.02 \times 10^{23} \text{ moles}$$

$$= 5.73 \times 10^{23} \text{ molecules}$$

As one molecule of HNO_3 contain atoms = 5 atoms

There fore,

5.73×10^{23} molecules contain No of atoms = $5 \times 5.73 \times 10^{23}$

$= 28.5 \times 10^{23}$

$= 2.87 \times 10^{24}$ atoms

Q.8 How many ions of Na^+ and Cl^{-1} will be present in 30 g of NaCl?

Given Data:

Given mass of NaCl = 30g

Molar mass of NaCl = 1(23)+3(35.5)

= 23 + 35.5

= 58.5g / mol

To Find:

No of Na+ ions = ?

No of Cl- ions = ?

Solution

$$\text{Number of formula units of NaCl} = \frac{\text{Given mass of NaCl}}{\text{Molar mass of NaCl}} \times N\frac{A}{A}$$

$$= \frac{30}{58.5} \times 6.02 \times 10^{23}$$

$$= 3.08 \times 10^{23} \text{ formula units}$$

As,

1 formula unit of NaCl contains number of Na^{+1} ions = 1

3.08×10^{23} formula units of NaCl contain number of Na^+ ions = 3.08×10^{23} ions

We knew that in NaCl

Number of Na^+ ions = number of Cl^- ion

Thus number of Cl ions = 3.08×10^{23}

Total number of sodium ions (Na^+) and chloride ions (Cl) = 6.16×10^{23} ions

Q.9 How many molecules of HCI will be required to have 10 grams of it?

Given mass of HCl = 10g

Molar mass of HCl = 1(1)+1(35.5)=1+35.5 = 36.5g / mol

Number of molecules of HCl = ?

Solution:

$$\text{Number of molecules of HCl} = \frac{\text{Given mass of HCl}}{\text{Molar mass of HCl}} \times N_A$$

$$= \frac{10}{36.5} \times 6.02 \times 10^{23}$$

$$\mathbf{= 1.64 \times 10^{23}\ molecules}$$

Q.10 How many grams of Mg will have the same number of atoms as 6 grams of C have?

Given data

Given mass of carbon = 6g

Atomic mass of carbon = 12 g/ mole

To find:

Mass of Mg = ?

Solution:

$$\text{No. of moles of carbon} = \frac{\text{Given mass of Carbon}}{\text{Molar mass of Carbon}}$$

$$= \frac{6}{12}$$

$$= 0.5\ \text{mol}$$

Number of carbon atoms = Number of moles $\times N_A$

Number of carbon atoms in 0.5 moles of carbon $= 0.5 \times 6.02 \times 10^{23}$ atoms

$= 3.01 \times 10^{23}$ atoms

As

The number of atoms of mg and carbon are same, so their number of moles are also equal.

Thus,

$$\text{No. of atoms of Mg} = \frac{\text{Given mass of Mg}}{\text{Molar mass of Mg}} \times N_A$$

$$3.01 \times 10^{23} = \frac{\text{mass of Mg}}{24} \times 6.02 \times 10^{23}$$

$$\text{Mass of Mg} = \frac{3.01 \times 10^{23} \times 24}{6.02 \times 10^{23}}$$

$$= 12g$$

So

12g of Mg will have same number of atoms as 6g of carbon.

Unit 2: Structure of Atoms
Exercise Questions

Exercise Multiple Choice Question Answers

1. **Which one of the following results in the discovery of proton**
   (a) cathode rays (b) canal rays (c) x-rays (d) alpha rays.
2. **Which one of the following is the most penetrating?**
   (a) protons (b) electrons. (c) neutrons (d) alpha particles
3. **The concept of orbit was used by**
   (a) J. J. Thomson (b) Rutherford (c) Bohr (d) Planck
4. **Which one of the following shell consists of three subshells.**
   (a) O-shell (b) N shell (c) L shell (d) M shell
5. **Which radioisotope is used for the diagnosis of tumor in the body?**
   (a) cobalt-60 (b) iodine-131 (c) strontium-90 (d) phosphorus-32
6. **When U-235 breaks up, it produces:**
   (a) electrons (b) neutrons (c) protons (d) nothing
7. **The p subshell has:**
   (a) one orbital (b) two orbitals (c) three orbitals (d) four orbitals
8. **Deuterium is used to make:**
   (a) light water (b) heavy water (c) soft water (d) hard water
9. **The isotope C-12 is present in abundance of:**
   (a) 96.9 % (b) 97.6 % (c) 99.7 % (d) none of these
10. **Who discovered the proton:**
    (a) Goldstein (b) J. T. Thomson (c) Neil Bohr (d) Rutherford

ANSWR KEY

1	b	3	c	5	a	7	c	9	d
2	c	4	d	6	b	8	b	10	a

Exercise Short Question Answers

Q.1 What is the nature of charge on cathode rays?

Ans: Cathode rays are negatively charged particles. J.J. Thomson discovered the e/m (charge/mass) ratio of cathode rays and found it equal to electron.

Q.2 Give five characteristics of cathode rays.

Ans: The characteristics of cathode rays are as under:

i. These rays travel in a straight line perpendicular to the cathode surface.
ii. They raise the temperature of the body on which they fall.
iii. Light is produced when these rays hit the sides of discharge tube.
iv. They can cast a sharp shadow of an opaque object if placed in their path.
v. The nature of rays does not depend upon the nature of as used in discharge tube.

Q.3 The atomic symbol of a phosphorus ion is given as $\left({}^{31}_{15}P^{3-}\right)$

(a) How many protons, electrons and neutrons are there in the ion?
(b) What is name of the ion?
(c) Draw the electronic configuration of the ion.
(d) Name the noble gas which has the same electronic configuration as the phosphorus ion has.

Ans:

a. **In $\left[{}^{31}_{15}P^{3-}\right]$ ion:**

i. Number of protons = 15
ii. Number of electron = 15+3=18 (P^{3-} has three more electrons than neutral P-atom)
iii. Number of neutrons = 31-15=16

b. The name of ion is Phosphide ion

c. Electronic configuration of ${}^{31}_{15}P^{-3}$ = $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$ (P^{3-} has three more electrons than neutral P-atom)

d. Argon has same electronic configuration as the phosphorous ion has.

Q.4 Differentiate between shell and subshell with examples of each.

Ans:

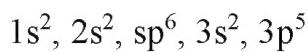
Shell	Sub-shell
i. *The circular path of an electron around the nucleus is called shell or principal energy level is called a shell.*	***i.*** *Each shell consists of smaller paths called subshells.*
ii. The shells are subdivided into subshells.	**ii.** The subshells are further composed of atomic orbitals.
iii. These are represented by K, L, M, N etc.	**iii.** Example: s, p, d and f are considered as the subshells of a shell. These are represented by s, p, d, f.

Q.5 An element has an atomic number 17. How many electrons are present in K, L and M shells of the atom?

Ans: Atomic number of element = number of electrons = 17
Therefore, its electronic configuration will be

K	L	M
2	8	7

OR

$1s^2, 2s^2, sp^6, 3s^2, 3p^5$

Q.6 Write down the electronic configuration of Al^{3+}. How many electrons are' present in its outermost shell?

Ans: Atomic number of Al = 13

Number of electrons of Al = 13

Number of electrons of Al^{3+} = 13-3 = 10 electrons.

Thus electronic configuration of Al^{3+} ion

K L

2 8

In terms of subshell: $1s^2, 2s^2, sp^6$

Therefore,

Number of electrons present in outer most shell of Al^{3+} = 8 electrons

Q.7 Magnesium has electronic configuration 2, 8, 2,

(a) How many electrons are in the outermost shell?

(b) In which subshell of the outermost shell electrons are present?

(c) Why magnesium tend to lose electrons?

Ans:

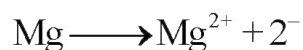
a. Electronic configuration of Mg = K L M
2 8 2

[Mg] = $1s^2, 2s^2, 2p^6, \mathbf{3s^2}$

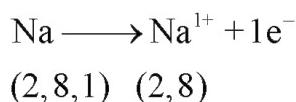
It has two electrons in the outermost shell.

b. The outermost electrons are present in "s" subshell of the 3rd shell (M).

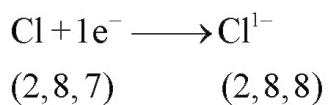
c. Magnesium is electropositive in character. It has the ability to lose its two electrons from its outermost shell.

$$Mg \longrightarrow Mg^{2+} + 2^-$$

Q.8 What will be the nature of charge on an atom when it loses an electron or when it gains an electron?

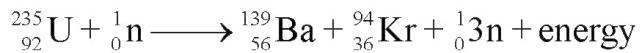
Ans: When an atom loses an electron, it acquires positive charge due to more number of protons in the nucleus e.g.

$$\underset{(2,8,1)}{Na} \longrightarrow \underset{(2,8)}{Na^{1+}} + 1e^-$$

When an atom gains an electron, it possesses negative charge due to more electrons

than protons in the atom e.g.

$$Cl + 1e^{-} \longrightarrow Cl^{1-}$$
$$(2,8,7) \qquad\qquad (2,8,8)$$

Q.9 For what purpose is U-235 used?

Ans: Radioactive isotope U-235 is used to generate electricity.

$$^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{139}_{56}Ba + ^{94}_{36}Kr + ^{1}_{0}3n + energy$$

Q.10 A patient has goiter, how will it be detected?

Ans: Isotopes of iodine-131 are used for diagnosis of goiter in thyroid gland. These radioactive isotopes are used as tracers in medicine to diagnose the presence of tumor in the human body.

Q.11 Give three properties of positive rays.

Ans: Positive rays are also called "canal rays". Their properties are:

i. These rays travel in straight line in a direction opposite to the cathode rays.

ii. These are positively charged rays.

iii. Mass of these particles was found equal to that of a proton or simple multiple of it.

Q.12 What are the defects of Rutherford's atomic model?

Ans: Rutherford's atomic model had following defects:

i. Stability of atom:

According to classical theory of radiations, electrons being charged particles should release or emit energy continuously. They should ultimately fall into the nucleus.

ii. Nature of spectrum:

If the electrons emit energy continuously they should form a continuous atomic spectrum but in fact, line atomic spectrum was observed.

Q.13 As long as electron remains in an orbit, it does not emit or absorb energy. When does it emit or absorb energy?

Ans: Electrons do not emit or absorb energy till they remain in their orbits. Electron emits energy when it jumps from high energy level to the lower energy level. An electron absorbs energy when it jumps from a lower energy orbit to a higher energy orbit. The change in energy is given by the following Planck's equation

$$AE = E_2 - E = hv \qquad \text{(Energy absorb)}$$

Where

E_1 = energy of lower energy orbit

E_2 = energy of higher energy orbit

"h" is Planck's" constant. Its value is 6.63×10^{-34} Js and frequency of light.

And

$E_2 - E_1 = -hv$ (Energy emitted)

Exercise Long Question Answers

Q.1 How are cathode rays produced? What are its five major characteristics?

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Q.2 How was it proved that electrons are fundamental particles of an atom?

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Q.3 Draw a labeled diagram to show the presence of protons in the discharge tube and explain how canal rays were produced.

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Q.4 How Rutherford discovered that atom has a nucleus located at the centre of the atom?

Ans: See Q. No. **5** (Subjective Part, Long Questions Answers)

Q.5 One of the postulates of Bohr's atomic model is that angular momentum of a moving electron is quantized. Explain its meaning and calculate the angular momentum of third orbit (i.e. n=3)

Ans:

Q.6 How did Bohr prove that an atom must exist?

Ans: See Q. No. **6** (Subjective Part, Long Questions Answers)

Q.7 What do you mean by electronic configuration? What are basic requirements while writing electronic configuration of an element (atom)?

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.8 Describe the electronic configuration of Na^+, Mg^{2+} and Al^{3+} ions. Do they have the same number of electrons in the outermost shell?

Ans:

(i) Na^+

Electronic configuration is shells =
K	L
2	8

In terms of subshell : is^2, $2s^2$, sp^6

(ii) Mg2+:

Electronic configuration in shell:
K	L
2	8

In terms of subshell: $1s^2$, $2s^2$, sp^6

(iii) Al3+

Electronic configuration in shell:

K	L
2	8

In terms of subshell: $1s^2$, $2s^2$, sp^6

Hence:

It is proved that all have 8 electrons in their outermost shells.

Q.9 Give the applications of isotopes in the field of radiotherapy and medicines.

Ans: See Q. No. **13** (Subjective Part, Long Questions Answers)

Q.10 What is an isotope? Describe the isotopes of hydrogen with diagrams.

Ans: See Q. No. **11** (Subjective Part, Long Questions Answers)

Unit 3 – Periodic Table and Periodicity of Properties

Exercise Questions

Exercise Multiple Choice Question Answers:

1. **The atomic radii of the elements in Periodic Table:**
   (a) Increase from left to right in a period (b) Increase from top to bottom in a group
   (c) Do not change from left to right in a period (d) Decrease from top to bottom in a group
2. **The amount of energy given out when an electron is added to an atom is called:**
   (a) Lattice energy (b) ionization energy (c) Electronegativity (d) Electron affinity
3. **Mendeleev Periodic Table was based upon the:**
   (a) Electronic configuration (b) atomic mass
   (c) Atomic number (d) completion of a subshell
4. **Long form of Periodic Table is constructed on the basis of:**
   (a) Mendeleev Postulate (b) Atomic number
   (c) Atomic mass (d) Mass number
5. **4th and 5th period of the long form of Periodic Table are called:**
   (a) Short periods (b) normal periods (c) Long periods (d) Very long periods
6. **Which one of the following halogen has lowest electronegativity?**
   (a) Flourine (b) chlorine (c) Bromine (d) Iodine

7. **Along the period, which one of the following decreases:**
   (a) Atomic radius (b) ionization energy (c) Electron affinity (d) Electronegativity
8. **Transition elements are:**
   (a) All gases (b) all metals (c) All non-metals (d) All metalloids
9. **Mark the incorrect statement about ionization energy:**
   (a) It is measured in kJmol^{-1} (b) It is absorption of energy
   (c) It decreases in a period (d) It decreases ina group
10. **Point out the incorrect statement about electron affinity:**
    (a) It is measured in klmol^{-1} (b) It involves release of energy
    (c) It decreases in a period (d) It decreases in a group

ANSWR KEY

1	b	3	b	5	c	7	a	9	c
2	d	4	b	6	d	8	b	10	c

Exercise Short Questions Answers

Q.1 Why noble gases are not reactive?

Ans: Noble gases are not reactive because they have their valence shells completely filled. They have 2 or 8 electrons in their valence shells. Their atoms do not have vacant spaces in their valence shell to accommodate more electrons. Therefore they do not gain, lose or share electrons.

Q.2 Why Cesium (at.no.55) requires little energy to release its one electron present in the outermost shell?

Cesium requires little energy because it has greater atomic size, more shielding effect (due to presence of more electrons) and low ionization energy due to which the hold of inner nucleus on valence.

Q.3 How is periodicity of properties dependent upon number of protons in an atom?

Ans: Number of protons in an atom represents atomic number of that element which increases regularly by one form element to element. So the arrangement of elements according to increasing atomic number shows the periodically in the electronic configuration of the elements that leads to periodicity in their properties.

Q.4 Why shielding effect of electrons makes cation formation easy?

Ans: The shielding effect of electrons makes the cation formation easy because it reduces the nuclear pull on the outermost electrons and they are less tightly held by the nucleus and can easily be lost from the outermost shell.

Q.5 What is the difference between Mendeleev's periodic law and modem periodic law?

Ans:

Mendeleevs periodic law	Modern periodic law
Properties of the elements are periodic function of their atomic masses.	Properties of the elements are periodic function of their atomic numbers.
Atomic masses is less fundamental property and it is the basis of mendeleevs periodic law.	Atomic number is more fundamental property and it is the basis of modern periodic law.

Q.6 What do you mean by groups and periods in a Periodic Table?

Ans: The horizontal rows of elements in a periodic table are called periods. The vertical columns in a periods table are called group. There are 18 groups in the long form of the periodic table. They are studied from top to bottom.

Q.7 Why and how are elements arranged in 4th period?

Ans: The elements (Na, Mg, Al, Si, P, S, Cl and Ar) are arranged in the 4th period because they are all having four electronic shells and are arranged by increasing atomic number

from left to right the period.

Q.8 Why the size of atom does not decrease regularly in a period?

Ans: The size of atom does not decrease regularly in a period. This irregularity in the transition metals is due to the involvement of d orbital. It provides poor shielding effect.

Q.9 Give the trend of ionization energy in a period.

Ans: ionization energy increases from left to right in a period and decreases from top to bottom in a group.

Reason:

It is because the size of atoms reduces and valence electrons are held strongly by the electrostatic force of nucleus.

Exercise Long Question Answers

Q.1 Explain the contributions of Mendeleev for the arrangement of elements in a Periodic Table.

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Q.2 Show why in a 'period' the size of an atom decreases if one moves from left to right.

Ans: See Q. No. **11** (Subjective Part, Long Questions Answers)

Q.3 Describe the trends of electronegativity in a period and in a group.

Ans: See Q. No. **15** (Subjective Part, Long Questions Answers)

Q.4 Discuss the important features of modem Periodic Table.

Ans: See Q. No. 7 (Subjective Part, Long Questions Answers)

Q.5 What do you mean by blocks in a periodic table and why elements were placed in blocks?

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.6 Discuss in detail the periods in Periodic Table?

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.7 Why and how elements are arranged in a Periodic Table?

Ans: See Q. No. **5** (Subjective Part, Long Questions Answers)

Q.8 What is ionization energy? Describe its trend in the Periodic Table?

Ans: See Q. No. **13** (Subjective Part, Long Questions Answers)

Q.9 Define electron affinity, why it increases in a period and decreases in a group in the Periodic Table.

Ans: See Q. No. **14** (Subjective Part, Long Questions Answers)

Q.10 Justify the statement, bigger size atoms have low ionization energy and have more shielding effect.

Ans: **Ionization Energy:**

The amount of energy required to remove the most loosely bound electron from the valence shell of an isolated gaseous atom is called ionization energy."

Shielding Effect:

"The decrease in the attractive force exerted by the nucleus on the valence shell electrons due to the presence of the electrons lying between the nucleus and valence shell is called shielding effect."

As we move down the group more and more shells lie between the valence shell and the nucleus of the atom, these additional shells reduce the electrostatic force felt by the electron present in the outermost shell which results more shielding effect by such bigger size atoms. Resultantly the valence shell electrons can be released easily. Therefore bigger size atoms have more shielding effect and low ionization energies.

Unit 4: Structure of Molecules

Exercise Questions

Exercise Multiple Choice Question Answer

1. **Atoms react with each other because:**
   (a) They are attracted to each other. (b) They are short of electrons
   (c) They want to attain stability (d) They want to disperse
2. **An atom having six electrons in its valence shell will achieve noble gas electronic configuration by:**
   (a) Gaining one electron (b) Losing all electrons
   (c) Gaining two electrons (d) Losing two electrons
1. **Considering the electronic configuration of atoms which atom with the given atomic number will be the most stable one?**
   (a) 6 (b) 8 (c) 10 (d) 12
3. **Octet rule is:**
   (a) Description of eight electrons (b) Picture of electronic configuration
   (c) Pattern of electronic configuration (d) Attaining of eight electrons
4. **Transfer of electrons between atoms results in:**
   (a) Metallic bonding (b) Ionic bonding
   (c) Covalent bonding (d) Coordinate covalent bonding

1. **When an electronegative element combines with electropositive element the type of bonding is:**
   (a) Covalent (b) Ionic
   (c) Polar covalent (d) Coordinate covalent
5. **A bond formed between two non-metals is expected to be:**
   (a) Covalent (b) ionic (c) Coordinate covalent (d) Metallic
6. **A bond pair in covalent molecules usually has:**
   (a) One electron (b) two electrons (c) Three electrons (d) Four electrons
7. **Which of the following compounds is not directional in its bonding?**
   (a) C (b) KBr (c) CO_2 (d) H_2O
8. **Ice floats on water because:**
   (a) Ice is denser than water (c) Water is denser than ice
   (b) Ice is crystalline in nature (d) Water molecules move randomly
9. **Covalent bond involves the**
   (a) Donation of electrons (b) acceptance of electrons
   (c) Sharing of electrons (d) repulsion of electrons
10. **How many covalent bonds does C_2H_2 molecule have?**
    (a) Two (b) Three (c) Four (d) Five
11. **Triple covalent bond involves how many number of electrons?**

(a) Eight (b) Six (c) Four (d) Only three

12. Which pair of the molecules has same type of covalent bonds?

(a) O_2 and HCl (b) O_2 and N_2 (c) O_2 and C_2 (d) O_2 and C_2H_2

13. Identify the compound which is not soluble in water.

(a) C_6H_6 (b) NaCl (c) KBr (d) $MgCl_2$

14. Which one of the following is an electron deficient molecule?

(a) NH_3 (b) BF_3 (c) N_2 (d) O_2

15. Identify which pair has polar covalent bonds.

(a) O_2 and Cl_2 (b) H_2O and N_2 (c) H_2O and C_2H_2 (d) H_2O and HCl

16. Which one of the following is the weakest force among the atoms?

(a) ionic force (b) metallic force (c) intermolecular force (d) covalent force

ANSWR KEY

1	c	4	d	7	a	10	c	13	b	16	b
2	c	5	b	8	b	11	c	14	d	17	d
3	c	6	b	9	b	12	d	15	a	18	c

Exercise Short Question Answers

Q.1 Why do atoms react?

Ans: Atoms react to form chemical bonds in order to get stability. Atoms achieve stability by attaining electronic configuration of inert gases by losing, gaining or sharing of electron.

Q.2 Why is the bond between an electropositive and an electronegative atom ionic in nature?

Ans: The bond between an electropositive and an electronegative atom is ionic in nature because electropositive atom due to low I.E. can lose electron easily and forms a positive ion whereas electronegative atom due to high electron affinity will accept that electron easily and forms a negative ion. In this way positive and negative ions are attracted by electrostatic force of attraction to form ionic bond.

Q.3 Ionic compounds are solids. Justify.

Ans: Ionic compounds are solids because they have strong electrostatic forces of attraction between positively and negatively charged ions which hold them in a three dimensional crystalline or solid form.

Example:

Potassium chloride (KCl) is a crystalline solid.

Q.4 More electronegative elements can form bonds between themselves. Justify.

Ans: More electronegative elements have high values of ionization energy and do not lose electrons. They share electrons between their own atoms to complete their valence shells and form covalent bond.

Q.5 Metals are good conductor of electricity. Why?

Ans: Metals are good conductors of electricity due to presence of mobile or free electrons.

Q.6 Ionic compounds conduct electricity in solution or molten form. Why?

Ans: Ionic compounds conduct electricity in solution or molten form because in these two states ionic compounds have free ions in them. When these free ions move in solution or molten state they become conductor of electricity.

Q.7 What type of covalent bond is formed in nitrogen molecule?

Ans: Triple covalent bond is formed in nitrogen molecule. In nitrogen molecule three bond pairs are involved in bond formation.

$$:N\vdots + \vdots N \longrightarrow :N\vdots\vdots N \longrightarrow N \equiv N \text{ or } N_2$$

Q.8 Differentiate between lone pair and bond pair of electron.

Ans.

Bonded pair	Lone pair
i. Bond pair of electrons is involve in bond formation	**i.** Lone pair of electron is not involved in bond formation.
ii. Electrons of bond pair are contributed by two atoms.	**ii.** Electrons of lone pair are contributed by one atom only.
iii. It is under the influence of two nuclei (atoms)	**iii.** It is under the influence of only one nucleus.
In a ammonia molecule there are three bond pairs of electrons	

Q.9 Describe at least two necessary conditions for the formation of a covalent bond.

Ans: Necessary conditions:

- **a.** Elements should be electronegative in nature.
- **b.** Electronegativity difference between bonding atoms should be very small or zero.
- **c.** The elements should share the electrons mutually.

d. There should be 4 or more valance electrons.

e. The ionization energies of the elements must be high.

Example: HCl, Cl_2, C_6H_6 and C_2H_2

Q.10 Why HCl has dipole-dipole forces of attraction?

Ans: HCl forms a polar covalent bond atoms due to difference of electro negativity between bonded atoms. There exists a dipole in the molecule. The positive end of one molecule attracts the negative end of there molecule. Hence dipole force. (Intermolecular forces) exist between HCl molecules.

Example:

$$H^{\delta+} — Cl^{\delta-} \text{------} H^{\delta+} — Cl^{\delta-}$$

Q.11 What is a triple covalent bond, explain with an example?

Ans: When each bonded atom contributes three electrons, three bond pairs are involved in bond formation. This type of bond is called triple covalent bond.

Representation:

It is represented by $(\equiv)$.

Example:

Triple covalent bond is formed in nitrogen molecule. In nitrogen molecule three bond pairs are involved in bond formation.

$$:N\vdots + \vdots N \longrightarrow :N\vdots\vdots N \longrightarrow N \equiv N \text{ or } N_2$$

Q.12 What is difference between polar and non-polar covalent bonds, explain with one example of each?

Ans: Difference between polar and non polar covalent

Polar Covalent Bond	Non Polar Covalent Bond
i. *It is a bond formed between two different types of atoms (hetero atoms).*	***i.*** *It is a bond formed between two similar atoms (homo atoms).*
ii. The shared pair of electrons is attracted by both the atoms un equally.	**ii.** The shared pair of electrons is attracted by both the atoms equally.
Examples. HCl, HBr, HF, H_2O etc	Examples. H_2, Cl_2, N_2, O_2 etc

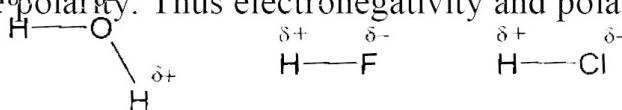
Q.13 Why a covalent bond becomes polar?

Ans: When there is a difference of electronegativity between two covalently bonded atoms, there will be unequal attraction for the bond pair of electrons between such atoms. It will result in the formation of polar covalent bond.

Examples: HCl, H_2O etc.

Q.14 What is relationship between electronegativity and polarity?

Ans: The polarity of a covalent bond depends upon the electronegativity difference between the bonded atoms. Higher the electronegativity difference between bonded atoms, greater will be the polarity. Thus electronegativity and polarity are directly related:

Examples: $\overset{\delta+}{H}-\overset{\delta-}{O}-\overset{\delta+}{H}$ $\overset{\delta+}{H}-\overset{\delta-}{F}$ $\overset{\delta+}{H}-\overset{\delta-}{Cl}$

Q.15 Why does ice float on water?

Ans: Ice floats on water because density of ice ($0.917g/cm^3$) is less than that of liquid water ($1.00g/cm^2$) at 0°C.

Q.16 Give the characteristic properties of ionic compounds.

Ans: Characteristics properties of ionic compounds.

i. Ionic compound are mostly crystalline solids.

ii. Ionic compounds are good conductors in solution and in molten form due to presence of free ions in them.

iii. Ionic compounds have high melting and boiling points. For example NaCl has melting point 800^0C and boiling point $1413^{\circ}C$.

iv. Ionic compounds dissolve in polar solvents e.g. NaCl dissolves in water.

Q.17 What characteristic properties do the covalent compounds have?

Ans: Characteristic properties of covalent compounds:

i. Melting boiling points: They have usually low melting and boiling point.

ii. Electrical conductivity: They are usually bad conductors of electricity. Polar compounds are conductors in their solutions in polar solvents.

iii. Solubility: They are usually insoluble in water but soluble in non-aqueous solvents like benzene, ether, alcohol and acetone.

iv. Crystal formation: Bigger molecules with three dimensional bonding form covalent crystals which are very stable and hard. They have high melting and boiling points.

Exercise Long Question Answers

Q.1 What is an ionic bond? Discuss the formation of ionic bond between sodium and chlorine atoms?

Ans: See Q. No. **4** (Subjective Part, Long Questions Answers)

Q.2 How can you justify that bond strength in polar covalent compounds is comparable to that of ionic compound?

Ans: See Q. No. **7** (Subjective Part, Long Questions Answers)

Q.3 What type of covalent bonds are formed between hydrogen, oxygen and nitrogen? Explain their bonding with dot and cross model.

Ans: See Q. No. **7** (Subjective Part, Long Questions Answers)

Q.4 How a covalent bond develops ionic character in it? Explain.

Ans:

Q.5 Explain the types of covalent bonds with at least one example of each type.

Ans: See Q. No. **5** (Subjective Part, Long Questions Answers)

Q.6 How a coordinate covalent bond is formed? Explain with examples?

Ans: See Q. No. **6** (Subjective Part, Long Questions Answers)

Q.7 What is metallic bonds? Explain the metallic bonding with the help of a diagram.

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.8 Define hydrogen bonding. Explain that how these forces affect the physical properties of compounds.

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.9 What are intermolecular forces? Compare these forces with chemical bond forces with reference to HCl molecule?

Ans:

Q.10 What is a chemical bond and why do atoms form a chemical bond?

Ans: See Q. No. **1** (Subjective Part, Long Questions Answers)

Q.11 What is octet rule? Why do atoms always struggle to attaint be nearest noble gas electronic configuration?

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Unit 5: Physical States of Matter

Exercise Questions

Exercise Multiple Choice Question Answers

1. **How many times liquids are denser than gases?**
(a) 100 times (b) 1000 times (c) 10,000 times (d) 100,000 times
2. **Gases are the lightest form of matter and their densities are expressed in terms of:**
(a) mg cm^{-3} (b) g cm^{-3} (c) g dm^{-3} (d) kg dm^{-3}
3. **At freezing point which one of the following coexists in dynamic equilibrium:**
(a) Gas and solid (b) liquid and gas (c) liquid and solid (d) all of these.
4. **Solid particles possess which one of the following motions?**
(a) Rotational motions (b) vibrational motions
(c) Translational motions (d) both translational and vibrational motions
5. **Which one of the following is not amorphous?**
(a) Rubber (b) plastic (c) glass (d) glucose.
6. **One atmospheric pressure is equal to how many Pascals:**
(a) 101325 (b) 10325 (c) 106075 (d) 10523
7. **In the evaporation process, liquid molecules which leave the surface of the liquid have:**
(a) Very low energy (b) moderate energy (c) very high energy (d) none of these
8. **Which one of the following gas diffuses faster?**
(a) Hydrogen (b) helium (c) fluorine (d) chlorine
9. **Which one of the following does not affect the boiling point?**
(a) Intermolecular forces (b) external pressure
(c) Nature of liquid (d) initial temperature of liquid
10. **Density of a gas increases, when its:**
(a) Temperature is increased (b) pressure is increased
(c) Volume is kept constant (d) none of these
11. **The vapour pressure of a liquid increases with the:**
(a) Increase of pressure (b) increase of temperature
(c) Increase of intermolecular forces (d) increase of polarity of molecules

ANSWR KEY

1	a	4	b	7	c	10	b
2	c	5	d	8	a	11	b
3	c	6	a	9	d		

Exercise Short Question Answers

Q.1 What is diffusion, explain with an example?

Ans: *The spontaneous mixing of particles of a substance by random motion and collisions, to form a homogeneous mixture is called diffusion.*

OR

Movement of molecules of a substance from the region of higher concentration to the region of lower concentration is called diffusion.

Example: When a few drops of ink are added in beaker of water, ink molecules move around and after a while spread in whole of the beaker. Thus diffusion has taken place.

Q.2 Define standard atmospheric pressure. What are its units? How it is related to Pascal?

Ans: **Standard atmospheric pressure:**

It is the pressure exerted by the atmosphere at the sea level. *"It is defined as **the pressure exerted by a mercury column of 760mm height at sea level"**. It is sufficient pressure to support a column of mercury 760mm in height at sea level.*

Units:

i. One atmosphere (1 atm): 1 atm is called standard pressure

ii. One pascal (1 Pa)

$1atm = 760mmHg = 760torr = 101325Nm^{-2} = 101325Pa$

(as $1mmHg = 1torr$

$1Nm^{-2} = 1Pa$)

Q.3 Why are the densities of gases lower than that of liquids?

Ans: Gases have lower densities than densities of liquids. It is due to the light mass and more volume occupied by the gases. Another reason for lower densities of gases is negligible intermolecular forces among the gases molecules. On the other hand liquid molecules are closely spaced and have strong intermolecular forces.

Q.4 What do you mean by evaporation, how it is affected by surface area?

Ans: **Evaporation:**

"*The process of changing of a liquid into a gas phase is called evaporation.*"

Affect of surface area on evaporation:

Evaporation is a surface phenomenon. Greater is surface area, greater is evaporation and vice versa.

Q.5 Define the term allotropy with examples.

Ans: Allotropy:

"The existence of an element in more than one forms, in same physical state is called allotropy."

Examples:

i. Oxygen has two allotropic forms i.e. oxygen (O_2) and ozone (O_3).

ii. Three allotropic forms of carbon are: Diamond, graphite and bucky balls.

Q.6 In which form sulphur exists at 100°C?

Ans: Sulphur exists in monoclinic form at 100°C

Q.7 What is the relationship between evaporation and boiling point of a liquid?

Ans: Relationship between evaporation and boiling point:

If the boiling point of a liquid is high, its evaporations slow. Because intermolecular forces are high in the liquid which have high boiling points. If boiling point is low then evaporation is high.

Exercise Long Question Answers

Q.1 Define Boyle's law and verify it with an example.

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Q.2 Define and explain Charles law of gases.

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Q.3 What is vapour pressure and how it is affected by intermolecular forces.

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.4 Define boiling point and also explain, how it is affected by different factors.

Ans: See Q. No. **10** (Subjective Part, Long Questions Answers)

Q.5 Describe the phenomenon of diffusion in liquids along with factors which influence it.

Ans: See Q. No. **12** (Subjective Part, Long Questions Answers)

Q.6 Differentiate between crystalline and amorphous solids.

Ans: See Q. No. **15** (Subjective Part, Long Questions Answers)

Exercise Solved Numericals

Q.1 Convert the following units:

a. 850 mm Hg to atm

b. 205000 Pa to atm

c. 560 torr to cm Hg

d. 1.25 atm to Pa

Solution:

a. 850 mmHg to atm

$$760mmHg = 1atm$$
$$1mmhg = \frac{1}{760} atm$$
$$850mmHg = \frac{1}{760} \times 850 atm$$
$$= 1.12atm$$

b. 205000 Pa to atm

$$101325Pa = 1atm$$
$$1Pa = \frac{1}{101325} atm$$
$$205000Pa = \frac{1}{101325} \times 205000 \text{ atm}$$
$$= 2.02 \text{ atm}$$

c. 560 torr to cm Hg

$$760 \text{ torr} = 760 \text{ mmHg}$$
$$= 76 \text{ cmHg}$$
$$1 \text{ torr} = \frac{76}{760} \text{cmHg}$$
$$560 \text{ torr} = \frac{76}{760} \times 560 cmHg$$
$$= 56cmHg$$

d. 1.25 atm to Pa

$$1 \text{ atm} = 101325Pa$$
$$1.25 \text{ atm} = 1.25 \times 101325Pa$$
$$= 126656 \text{ Pa}$$

Q.2 <u>Convert the following units.</u>

a. 750°C to K **b. 150°C to K** **c. 100 K to °C** **d. 172 K to °C**

Solution:

a. 750°C to K

$$T (°C) = 750 °C$$
$$T (K) = ?$$
$$T (K) = T(°C) + 273$$
$$= 750 + 273$$
$$= 1023K$$

b. 150ºC to K

$$T(°C) = 150\ °C$$
$$T(K) = ?$$
$$T(K) = T(°C) + 273$$
$$= 150 + 273$$
$$= 423$$

c. 100 K to °C

$$T(K) = 100\ K$$
$$T(°C) = ?$$
$$T(°C) = T(K) - 273.15$$
$$= 100 - 273$$
$$= -173°C$$

d. 172 K to ºC

$$T(K) = 172\ K$$
$$T(°C) = ?$$
$$T(°C) = T(K) - 273$$
$$= 172 - 273$$
$$= -101°C$$

Q.3 <u>A gas at pressure 912 mm of Hg has volume 450cm^3. What will be its volume at 0.4 atm.</u>

Given Data:

$$P_1 = 912\ mm\ Hg = \frac{912\ mm\ Hg}{760\ mm\ Hg}$$
$$= 1.2\ atm$$
$$V_1 = 450\ cm^3$$
$$P_2 = 0.4\ atm$$

Required:

$$V_2 = ?$$

Using the equation of Boyle's Law:

$$P_1V_1 = P_2V_2$$

Solution:

By putting the values:

$$1.2\ atm \times 450\ cm^3 = 0.4\ atm \times V_2$$
$$V_2 = \frac{1.2\ atm \times 450\ cm^3}{04\ atm}$$
$$V_2 = \frac{12}{4} \times 450\ cm^3$$
$$V_2 = 3 \times 450\ cm^3$$
$$\mathbf{V2 = 1350\ cm^3}$$

Q.4 A gas occupies a volume of 800 cm^3 at 1 atm, when it is allowed to expand up to 1200 cm^3 what will be its pressure in mm of Hg.

Given Data:

$$P_1 = 1 \text{ atm}$$
$$V_1 = 800 \text{ cm}^3$$
$$V_2 = 1200 \text{ cm}^3$$

Required:

$$P_2 = ?$$

Using the equation of Boyle's Law:

$$P_1V_1 = P_2V_2$$

Solution:

By putting the values

$$1 \text{ atm} \times 800 \text{ cm}^3 = P_2 \times 1200 \text{ cm}^3$$

$$P_2 = \frac{1 \text{ atm} \times 800 \text{ cm}^3}{1200 \text{ cm}^3}$$

$$P_2 = \frac{2}{3} \text{ cm}^3$$

$$P_2 = 0.667 \text{ atm}$$

As $\quad 1 \text{ atm} = 760\text{mmHg}$

So $\quad 0.66 \text{ atm} = 760 \times 0.66\text{mmHg}$

$$= 506.66\text{mmHg}$$

Q.5 It is desired to increase the volume of a fixed amount of gas from 87.5 to 118 cm^3 while holding the pressure constant. What would be the final temperature if the "initial temperature is 23°C.

Given Data:

$$V_1 = 87.5 \text{ cm}^3$$
$$V_2 = 118 \text{ cm}^3$$
$$T_1 = 23°C\ (23+273)\ K = 296K$$

Required:

$$T_2 = ?$$

By using the equation of charle's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Solution

$$T_2V_1 = V_2 \times T_1$$

Or

$$T_2 = \frac{V_2\,'T_1}{V_1}$$

By putting the values

$$T_2 = \frac{118\text{cm}^3 \times 296K}{87.5 \text{ cm}^3}$$

$$T_2 = 399K$$

T_2 can be converted into Celsius scale as:

$$T_2 = 299 - 273 = 126°C$$

Q.6 A sample of gas is cooled at constant pressure from 30°C to 10°C. Comment:

a. Will the volume of the gas decrease to one third of its original volume?

b. If not, then by what ratio will the volume decrease?

Solution:

a.

$T_1 = 30°C = (30+273)\ K = 303K$

$T_2 = 10°C = (10+273)K = 283K$

$V_1 = 1\ dm^3$

Required: $V_2 = ?$

Solution:

By using the equation of Charle's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = \frac{V_1}{T} \times T_2$$

By putting the values

$$= \frac{1dm^3}{303K} \times 283K$$

$$V_2 = 0.93dm^3$$

The volume of gas will not decrease to one third of its original volume.

(b) The volume decreases in the ration 1:0.93.

Q.7 A balloon that contains 1.6 dm³ of air at standard temperature and pressure is taken under water to a depth at which its pressure increases to 3.0 atm. Suppose that temperature remain unchanged, what would be the new volume of the balloon. Does it contract or expand?

Given Data:

$P_1 = 1\ atm$

$V_1 = 1.6\ dm^3$

$P_2 = 3.0\ atm$

Required:

$V_2 = ?$

Solution:

By using the equation of Boyle's law

$$P_1V_1 = P_2V_2$$

By putting the values

$$1\ atm \times 1.6\ dm = 3\ atm \times V_2$$

$$V_2 = \frac{1\ atm \times 1.6\ dm^3}{3\ atm}$$

$$V_2 = 0.53\ dm^3$$

The new volume of balloon is $0.55dm^3$. It will contract.

Q.8 A sample of neon gas occupies 75.0 cm^3 at very low pressure of 0.4 atm. Assuming temperature remain constant what would be the volume at 1.0 atm pressure?

Given Data:

$$P_1 = 0.4 \text{ atm}$$
$$V_1 = 75.0 \text{ cm}^3$$
$$P_2 = 1 \text{ atm}$$

Required:

$$V_2 = ?$$

Solution

By using the equation of Boyle's law

$$P_1V_1 = P_2V_2$$

By putting the values

$$0.4 \text{ atm} \times 75 \text{ cm}^3 = 1 \text{ atm} \times V_2$$

$$V_2 = \frac{0.4 \text{ atm} \times 75 \text{ cm}^3}{1 \text{ atm}}$$

$$\mathbf{V_2 = 30 \text{ cm}^3}$$

Q.9 A gas occupies a volume of 35.0 dm^3 at 17°C. If the gas temperature rises to 34°C at constant pressure, would you expect the volume to double? If not calculate the new volume.

Given Data:

$$T_1 = 17\ °C$$
$$= 273 + 17 = 290 \text{ K}$$

$$V_1 = 35 \text{ dm}^3$$
$$T_2 = 34\ °C$$
$$= 273 + 34 = 307 \text{ K}$$

Required:

$$V_2 = ?$$

Solution:

Volume will not be doubled because the absolute temperature is not doubled.

By using the equation of Charle's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

By putting the values

$$\frac{35 \text{ dm}^3}{290 \text{ K}} = \frac{V_2}{307 \text{ K}} \text{ or}$$

$$V_2 = \frac{35 \text{ dm}^3 \times 307 \text{ K}}{290 \text{ K}}$$

$$37 \text{ dm}^3 = V_2$$

Q.9 The largest moon of Saturn, is Titan. It has atmospheric pressure of 1.6×10^5 Pa. What is the atmospheric pressure in atm? Is it higher than earth's atmospheric pressure?

Solution:

We know that

$$1 \text{atm} = 101325 \text{ Pa}$$

Atmospheric pressure of titan in Pascal $= 1.6\times10^5$Pa.

Atmospheric pressure of titan in atm $= \dfrac{1.6\times10^5}{101325}$

$= 1.58$atm

Thus the atmosphere pressure of titan (1.58 atm) is greater than the atmospheric pressure of earth (1.0atm).

Unit 6: Solutions

Exercise Questions

Exercise Multiple Choice Question Answers

1. **Mist is an example of solution:**
   (a) Liquid in gas (b) Gas in liquid (c) Solid in gas (d) Gas in solid
2. **Which one of the following is a 'liquid in solid' solution?**
   (a) Sugar in water (b) Butter (c) Opal (d) Fog
3. **Concentration is ratio of:**
   (a) Solvent to solute (b) Solute to solution (c) Solvent to solution (d) Both 'A' and 'B'
4. **Which one of the following solutions contains more water?**
   (a) 2M (b) 1M (c) 0.5 M (d) 0.25 M
5. **A 5 percent (w/w) sugar solution means that:**
   (a) 5 g of sugar is dissolved in 90 g of water
   (b) 5 g of sugar is dissolved in 100 g of water
   (c) 5 g of sugar is dissolved in 105 g of water
   (d) 5 g of sugar is dissolved in 95 g of water
6. **If the solute-solute forces are strong enough than those of solute-solvent forces. The solute:**
   (a) Dissolves readily (b) Does not dissolve
   (c) Dissolves slowly (d) Dissolves-and precipitates.

7. **Which one of the following will show negligible effect of temperature on its solubility?**
   (a) KCl (c) $NaNO_3$ (b) KNO_3 (d) NaCl
8. **Which one of the following is heterogeneous mixture?**
   (a) Milk (b) ink (c) Milk of magnesia (d) Sugar solution
9. **Tyndall effect is shown by:**
   (a) Sugar solution (c) jelly (b) Paints (d) Chalk solution
10. **Tyndall effect is due to:**
    (a) Blockage of beam of light (b) Non-scattering of beam of light
    (c) Scattering of beam of light (d) Passing through beam of light
11. **If 10 cm^3 of alcohol is dissolved in 100 g of water, it is called:**
    (a) % w/w (b) %w/v (c) % v/w (d) %v/v
12. **When a saturated solution is diluted it turns into:**
    (a) Supersaturated solution (b) Saturated solution
    (c) A concentrated solution (d) Unsaturated solution
13. **Molarity is the number of moles of solute dissolved in:**
    (a) lkg of solution (b) 100 g of solvent (c) 1 dm^3 of solvent (d) 1 dm^3 of solution.

ANSWR KEY

1	a	3	b	5	b	7	d	9	c	11	c	13	d

2	b	4	d	6	c	8	c	10	c	12	d	

Exercise Short Question Answers

Q.1 Why suspensions and solutions do not show Tyndall effect, while colloids do?

Ans: Suspensions and solutions do not show Tyndall effect because in suspensions particles are so big that light is blocked and difficult to pass. But in solution particles are so small that they cannot scatter the rays of light, thus do not show Tyndall effect. But colloids can show Tyndall effect because particles scatter the path of light rays thus emitting the beam of light i.e., exhibit the Tyndall effect.

Q.2 What is the reason for the difference between solutions, colloids and suspensions?

Ans: The differentiation between solutions, colloids and suspensions is based upon the particle size. In colloidal solutions the particles size is intermediate between true solutions and suspensions.

Q.3 Why does not the suspension form a homogeneous mixture?

Ans: In suspension particles remain un-dissolved due to their big size. After sometime particles settle down under the action of gravity, therefore suspension does not forma homogeneous mixture.

Q.4 How will you test whether given solution is a colloidal solution or not?

Ans: We will pass light in the solution, if the given solution scattered the light then it is a colloidal solution. It solution does not scatter the light then it is not colloidal solution.

Q.5 Classify the following into true solution and colloidal solution Blood, starch solution, glucose solution, tooth paste, copper sulphate solution, silver nitrate solution.

Ans:

True Solutions	Colloidal Solution
Glucose solution, Copper sulphate solution, silver nitrate solution	Blood, tooth paste, starch solution

Q.6 Why we stir paints thoroughly before using?

Ans: Paints are heterogeneous mixture of un-dissolved particles in a given medium. Particles settle down after sometime. So we stir paints to mix thoroughly before using.

Q.7 Which of the following will scatter light and why? Sugar solution, soap solution and milk of magnesia.

Ans: **Soap solution:**

Soap solution will scatter light (Tyndall effect) because it is colloidal solution and its particles are large enough to scatter the light.

Sugar Solution:

Sugar solution will not scatter light because the particles of sugar solution are so small that they cannot scatter light.

Milk of Magnesia:

Milk of magnesia cannot scatter the light because it is suspension and its particles are so big that light is blocked.

Q.8 What do you mean by "like dissolves like?" Explain with examples.

Ans: "Like dissolves like" means that polar substances are dissolved in polar solvents and non polar substances are soluble in non polar solvents.

For example: NaCl (polar) dissolves in water (polar solvent) and does not dissolve in benzene (non polar)

Q.9 How does nature of attractive forces of solute-solute and solvent-solvent affect the solubility?

Ans: Solubility depends upon solute solvent attractions. If the attractive forces between solvent are stronger the solubility is greater. If the attractive forces become weaker in solute there will be greater solubility.

If the attractive forces between solute particles are stronger than solute solvent forces, solute remains insoluble and solution is not formed.

Q.10 How you can explain the solute-solvent interaction to prepare a NaCI solution?

Ans: When NaCl is added in water it dissolves readily because the attractive forces between the ions of NaCl and polar molecules of water are strong enough to overcome the attractive forces between Na^+ and Cl^- ions in solid NaCl crystal. In this process, positive end of the water dipole is oriented towards the Cl^- ions and the negative end of water dipole is oriented towards the Na^+ ions. These ion-dipole attractions between Na^+ ions and water molecules, Cl^- ions and water molecules are so strong that they pull these ions from their positions in the crystal and thus NaCl dissolves.

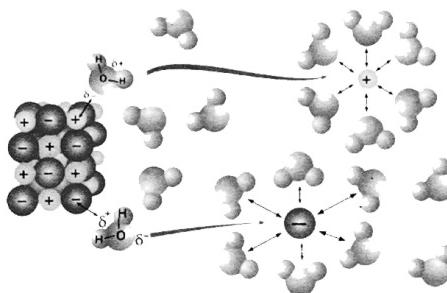



Fig. 6.2 Inter-action of solute and solvent to form solution.

Q.11 Justify with an example that solubility of a salt increases with the increase in temperature

Ans: Solubility of some salts which are usually ionic in nature increases with the increase in temperature for such solutes. It means that heat is required to break the attractive forces between the ions of solute. This process is called endothermic.

For example:

Solubility of KNO_3 and KCl can be enhanced by increasing temperature.

Q.12 What do you mean by volume/volume %?

Ans: It is the volume in cm^3 of a solute dissolved in 100 cm^3 of the solution.

For example:

30% of alcohol solution means 30 cm^3 of alcohol dissolved in sufficient amount of water, so that the total volume of the solution becomes 100 cm^3.

$$\frac{\%\text{Volume}}{\%\text{Volume}} = \frac{\text{Volume of solute}\left(cm^3\right)}{\text{Volume of solutio}\left(cm^3\right)} \times 100$$

Exercise Long Question Answers

Q.1 What is saturated solution and how it is prepared?

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Q.2 Differentiate between dilute and concentrated solutions with a common example.

Ans: See Q. No. **4** (Subjective Part, Long Questions Answers)

Q.3 Explain, how dilute solutions are prepared from concentrated solutions?

Ans: See Q. No. 7 (Subjective Part, Long Questions Answers)

Q.4 What is molarity and give its formula to prepare molar solution?

Ans: See Q. No. **6** (Subjective Part, Long Questions Answers)

Q.5 Explain the solute-solvent interaction for the preparation of solution.

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.6 What is general principle of solubility?

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.7 Discuss the effect of temperature on solubility.

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.8 Give the five characteristics of colloid.

Ans: See Q. No. **10** (Subjective Part, Long Questions Answers)

Q.9 Give at least five characteristics of suspension

Ans: See Q. No. **10** (Subjective Part, Long Questions Answers)

Exercise Solved Numerical

Q.1 A solution contains 50 g of sugar dissolved in 450 g of water. What is concentration of this solution?

Given Data:

Mass of sugar solute = 50g

Mass of water solvent = 450g

Required:

Concentration of solution (% m/m) = ?

Solution:

$$\% \text{ m/m} = \frac{\text{Mass of solute}(g)}{\text{Mass of solute}(g) + \text{Mass of solvent}(g)} \times 100$$

Solution:

$$\% \text{ m/m} = \frac{50g}{50g + 45g} \times 100$$

$$= \frac{50g}{500g} \times 100$$

Thus,

% m/m = 10% m/m

Q.2 If 60 cm^3 of alcohol is dissolved in 940 cm^3 of water, what is concentration of this solution?

Given Data:

Volume of alcohol solute = v = 60 cm^3

Volume of water solvent = v = 940 cm^3

Required Data:

Concentration of solution (% v/v) = ?

Formula:

$$\% \text{v/v} = \frac{\text{volume of solute}(cm^3)}{\text{volume of solute}(cm^3) + \text{volume of solvent}(cm^3)} \times 100$$

Solution:

$$\% \text{ v/v} = \frac{60 cm^3}{60 \text{ cm}^3 + 940 cm^3} \times 100$$

$$= \frac{60 \text{ cm}^3}{1000 cm^3} \times 100$$

Thus **% v/v = 6%v/v**

Q.3 How much salt will be required to prepare following solutions (atomic mass: K=39; Na=23; S=32; O=16 and H=I)

(a) 250 cm^3 of KOH solution of 0.5 M

(b) 600 cm^3 of $NaNO_3$ solution of 0.25 M

(c) 800 cm^3 of Na_2SO_4 solution of 1.0 M

Ans:

(a) 250cm^3 of KOH solution of 0.5M

Given Data:

Molarity of solution = (M) = 0.5 M

Volume of solution = 250 cm^3 = $\frac{250}{1000}$ dm^3=0.25dm^3

Molar mass of KOH = 39+16+1=56gmol^{-1}

Required Data:

Mass of KOH =?

Solution:

$$\text{Molarity}=\frac{\text{Mass of solute}(g)}{\text{Molar mass of solute }(gmol^{-1})\times\text{volume of solution }(dm^3)}$$

$$0.5M=\frac{\text{Mass of solute}(g)}{56g\ mol^{-1}\times 0.25dm^3}$$

$$\text{Mass of solute} = 0.5\times 56\times 0.25$$

=7g

(b) 600cm^3 of $NaNO_3$ solution of 0.25M

Given Data:

Molarity of $NaNO_3$ solutoin = (M) = 0.25M

Volume of solution = 600 cm^3 $=\frac{600}{1000}=0.6dm^3$

Molar mass of $NaNO_3$ = 23 + 14 + 3(16)

= 85gmol^{-1}

Required:

Amount of $NaNO_3$ = m =?

Solution:

Using the formula:

$$\text{Molarity}=\frac{\text{Mass of solute}(g)}{\text{Molarmass of solute}(gmol^{-1})\times\text{Volume of solution}(dm^3)}$$

$$\text{Molarity}=\frac{\text{Mass of solute}(g)}{85gmol^{-1}\times 0.6dm^3}$$

$$\text{Mass of solute} = 0.25\times 85\times 0.6$$

$$\text{Mass of solute} = 12.75g$$

(c) 800cm^3 of $Na_2\ SO_4$ solution of 1.0M

Given Data:

Molarity of Na_2SO_4 solution = M = 1 M

Volume of solution = V = 800 cm^3 $=\frac{800}{1000}=0.8dm^3$

Molecular mass of Na_2SO_4 = 2(23) + 32 +4(16)

= 46 + 32 + 64

= 142gmol^{-1}

Required:

Mass of Na_2SO_4 =?

Solution:

Using the formula

$$\text{Molarity}=\frac{\text{Mass of solute}(g)}{\text{Molar mass of solute}(gmol^{-1})\times\text{Volume of solution}(dm^3)}$$

$$1.0M=\frac{\text{Mass of solute}}{142gmol^{-1}\times 0.8dm^3}$$

$$\text{Mass of solute} = 1.0 \times 142 \times 0.8$$
$$= 113.6\text{g}$$

Q.4 When we dissolve 20 g of NaCl in 400 cm³ of solution, what will be its molarity?

Given Data:

Mass of NaCl = 20g

Molar mass of NaCl = 23 + 35.5 = 58.5gmol^{-1}

Volume of Solution = 400 cm^3 $= = \frac{400}{1000} 0.4\text{dm}^3$

Required:

Molarity of solution = ?

Solution:

Using the formula:

$$\text{Molarity} = \frac{\text{Mass of solute}(\text{g})}{\text{Molar mass of solute}(\text{gmol}^{-1}) \times \text{Volume of solution}(\text{dm}^3)}$$

$$= \frac{25\text{g}}{58.5\text{mol} \times 0.4(\text{dm}^3)}$$

$$= \frac{20}{23.4} = 0.85\text{M}$$

Q.5 We desire to prepare 100 cm³ 0.4 M solution of Mg Cl₂, how much Mg Cl₂ is needed?

Given Data:

Molarity of solution = 0.4 M

Volume of Solution = 100cm$^3 = \frac{100}{1000}\text{dm}^3 = 0.1\text{dm}^3$

Mass of $MgCl_2$ = 24 + 2(35.5) = 95g

=24+71=95gmol^{-1}

Required:

Mass of $MgCl_2$ = ?

Solution:

Using the formula:

$$\text{Molarity} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}(\text{gmol}^{-1}) \times \text{Volume of solutoin}(\text{dm}^3)}$$

$$0.4\text{M} = \frac{\text{Mass of solute}(\text{g})}{95\text{g mol}^{-1} \times 0.1\text{dm}^3}$$

Mass of solute $= 0.4 \times 95 \times 0.1$

$= 3.8\text{g}$

Q.6 12M H_2S0_4 solutions is available in the laboratory. We need only $500cm^3$ of 0.1 M solution, how it will be prepared?

Given Data:

Molarity of concentrated H_2SO_4 solution $= M_1 = 12$ M
Molarity of dilute H_2SO_4 solution $= M_2 = 0.1$ M
Volume of dilute H_2SO_4 solution $= V_2 = 500cm^3$

Required:

Volume of concentrated H_2SO_4 solution $= V_1 = ?$

Solution:

i. Determination of volume of concentrated solution:

Concentrated solution =Dilute solution

$$M_1 V_1 = M_2 V_2$$

$$12 \times V_1 = 0.1 \times 500$$

$$V_1 = \frac{0.1M \times 500\ cm^3}{12M}$$

Thus,

$$4.16\ cm^3$$

ii. Preparation of solution

We take $4.16cm^3$ of concentrated 12M H_2SO_4 solution with the help of graduated pipette

and put in a measuring flask of $500cm^3$. Add water upto the mark, present at the neck o

flask. Now it is 0.1 molar solution of H_2SO_4 .

Unit 7: Electrochemistry
Exercise Questions

Exercise Multiple Choice Question Answers:

1. **Spontaneous chemical reactions take place in:**
   (a) Electrolytic cell (b) Galvanic cell (c) Nelson's cell, (d) Down's cell
2. **Formation of water from hydrogen and oxygen is:**
   (a) Redox reaction (b) Acid-base reaction
   (c) Neutralization (d) Decomposition
3. **Which one of the following is not an electrolytic cell?**
   (a) Downs cell (b) Galvanic cell (c) Nelson's cell (d) Both a and c
4. **The oxidation number of chromium in $K_2Cr_2O_7$ is:**
   (a) +2 (b) +6 (c) +7 (d) +14
5. **Which one of the following is not an electrolyte?**
   (a) Sugar solution (b) Sulphuric acid solution
   (c) Lime solution (d) Sodium chloride solution
6. **The most common example of corrosion is:**
   (a) Chemical decay (b) Rusting of iron
   (c) Rusting of aluminium (d) Rusting of tin

7. **Nelson's cell is used to prepare caustic soda along with gases. Which of the following gas is produced at cathode:**
   (a) Cl_2 (b) H_2 (c) O_3 (d) O_2
8. **During the formation of water from hydrogen and oxygen, which of the following does not occur:**
   (a) Hydrogen has oxidized (b) Oxygen has reduced
   (c) Oxygen gains electrons (d) Hydrogen behaves as oxidizing agent
9. **The formula of rust is:**
   (a) $Fe_2O_3.nH_2O$ (b) Fe_2O_3 (c) $Fe(OH)_3.nH_2O$ (d) $Fe(OH)_3$
10. **In the redox reaction between Zn and HCl, the oxidizing agent is:**
    (a) Zn (b) H^+ (c) Cl^- (d) H

ANSWR KEY

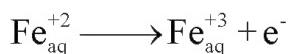
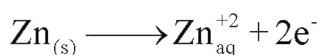
1	b	2	a	3	d	4	b	5	a	6	b	7	b
8	d	9	a	10	b								

Exercise Short Question Answers

Q.1 Define oxidation in terms of electrons. Give an example.

Ans: Oxidation is the loss of electron by an atom or an ion e.g,

$$Zn_{(s)} \longrightarrow Zn^{+2}_{aq} + 2e^-$$

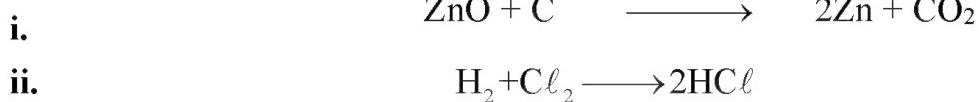
$$Fe^{+2}_{aq} \longrightarrow Fe^{+3}_{aq} + e^-$$

Q.2 Define reduction in terms of loss or gain of oxygen or hydrogen. Give an example.

Ans: Reduction:

"The addition of hydrogen or removal of oxygen during a chemical reaction."

Examples:

i. $ZnO + C \longrightarrow 2Zn + CO_2$

ii. $H_2 + C\ell_2 \longrightarrow 2HC\ell$

Q.3 What is difference between valency and oxidation state?

Ans:

Valency	Oxidation Number or state
• The combining capacity of an element with other element is called value of	• The apparent charge assigned to an atom of an element in a molecule or ion is called oxidation state.
• While assigning valency the sign is followed by the number i,e 2+	• No sign
• For example valency of sodium is 1+	• For example oxidation number of sodium is +1

Q.4 Differentiate between oxidizing and reducing agents

Ans:

Oxidizing agent	Reducing Agent
i. A species that oxidizes a substance by taking electrons from it, is called an oxidizing agent.	**i.** A species that reduces a substance by donating electrons to it is called reducing agent.
ii. Non metals are good oxidizing agents.	**ii.** Metals are good reducing agents.
iii. They are more electronegative in nature.	**iii.** They are more electropositive.
iv. Its oxidation number decreases.	**iv.** Its oxidation number decreases.
v. Examples	**vii.** Examples
vi. $S + O_2 \longrightarrow SO_2$	**viii.**

Q.5 Differentiate between strong and weak electrolytes.

Ans:

Strong electrolyte	Weak electrolyte
The electrolyte which ionize completely in solution and produce more ions, are called strong electrolyte.	*The electrolytes which ionize to a small extent when dissolve in water and could not produce more ions are called weak electrolytes.*
Examples: NaCl, NaOH, H_2SO_4 etc.	Examples: $Ca(OH)_2, CH_3COOH$ $CH_3COOH_\ell + H_2O_\ell \longrightarrow CH_3COO^-_{aq} + H_3O^+$

Q.6 How electroplating of tin on steel is carried out?

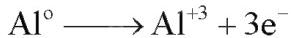
Ans: In electroplating of silver, when current is passed through the cell. A.g. ions present in the electrolyte solution migrate towards the cathode and deposit after picking up electrons. The anode consists of silver bar or sheet. Which is oxidized to Ag ions which dissolve in solution and migrate towards the cathode where they are discharged and deposited on the object

At anode: $Ag_{(s)} \longrightarrow Ag^+_{(aq)} + \bar{e}$

At cathode: $Ag^+_{(aq)} \longrightarrow Ag_{(s)}$

Q.7 Why steel is plated with nickel before the electroplating of chromium.

Ans: The steel is usually plated first with nickel or copper then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is stripped off.

Q.8 How can you explain, that following reaction is oxidation in terms of increase of oxidation number Al" $Al^o \longrightarrow Al^{+3} + 3e^-$

Ans: Increase in oxidation number is called oxidation oxidation number of Al in creases from zero to + 3 as given below $A\ell \longrightarrow A\ell^{+3} + 3e^-$

Q.9 How can you prove so it is an oxidation reaction with an example that conversion of an ion to an atom is an oxidation process?

Ans: Conversion of anion into an atom is an oxidation process.

Example:

When anions (negatively charged ions) lose electron, they are converted into atoms and oxidized.

$$Cl^- \xrightarrow{Oxidation} Cl + le^-$$

Q.10 Why does the anode carries negative charge in galvanic cell but positive charge in electrolytic cell? Justify with comments.

Ans: In Gavanic cell, electrons are lost by the atoms at anode plate which makes it electron efficient therefore it carries negative charge. In electrolytic cell, electrons are gained by cations from anode which makes it electron deficient therefore it carries positive charge.

Q.11 Where do the electrons flow from Zn electrode in Daniel's cell?

Ans: In Daniel cell, the electrons takes flow from Zn electrode (anode) towards the cathode made up of copper through the external circuit.

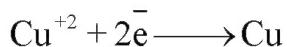
Q.12 Why do electrodes get their names 'anode' and cathode in galvanic cell?

Ans: In galvanic cell anode and cathode get their names depending upon the process taking place on them.

Anode: is an electrode where oxidation takes place

e.g. $$Zn \longrightarrow Zn^{+2} + 2e^-$$

Cathode: is an electrode where reduction takes place

$$Cu^{+2} + 2\bar{e} \longrightarrow Cu$$

In galvanic cell, oxidation takes place at anode while reduction takes place at cathode. And oxidation always takes place at anode while reduction always takes place at cathode.

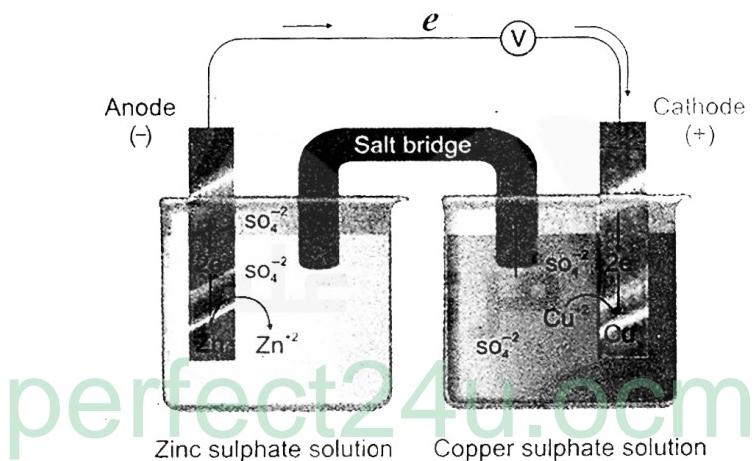
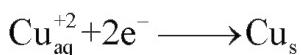



Fig. 7.3 A Daniel Cell

Q.13 What happens at the cathode in a galvanic cell?

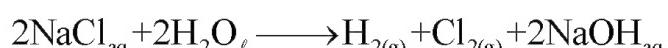
Ans: In galvanic cell, reduction takes place at the cathode as:

$$Cu^{+2}_{aq} + 2e^- \longrightarrow Cu_s$$

Q.14 Which solution is used as an electrolyte in Nelson's cell?

Ans: An (aqueous solution of NaCl called brine? is used as electrolyte in Nelson's cell.

Q.15 Name the by-products produced in Nelson's cell?

Ans: Hydrogen gas (H_2) and chlorine gas (Cl_2) are the by-product of Nelson's cell as

$$2NaCl_{aq} + 2H_2O_{\ell} \longrightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{aq}$$

Q.16 Why galvanizing is done?

Ans: The process of coating a thin layer of zinc on iron is called galvanizing. Galvanizing is done to protect the iron against corrosion even after the required coating surface is broken.

Q.17 Why an iron grill is painted frequently?

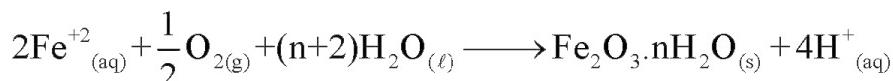
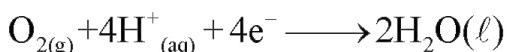
Ans: Iron grill is painted frequently to protect it from rusting. Paint layer proctect iron from attack of moisture and oxygen.

Q.18 Why O_2 is necessary for rusting?

Ans: O_2 is necessing for rusing because it acts as oxidizing agent. It accepts electrons from Fe which is covered to Fe+2 and then to Fe+3. Oxygen combines with Fe+3 to form rust (Fe_2O_3 H_2O)

The overall cell nraction for corrosion of ions is

$$O_{2(g)} + 4H^+_{(aq)} + 4e^- \longrightarrow 2H_2O(\ell)$$

$$2Fe^{+2}_{(aq)} + \frac{1}{2}O_{2(g)} + (n+2)H_2O_{(\ell)} \longrightarrow Fe_2O_3.nH_2O_{(s)} + 4H^+_{(aq)}$$

Q.19 In electroplating of chromium, which salt is used as an electrolyte?

Ans: Chromium sulphate with few drops of H_2SO_4 acts as electrolyte.

Q.20 Write the redox reaction taking place during the electroplating of chromium?

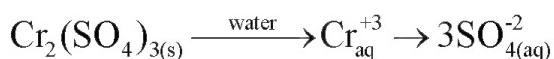
Ans: At anode:

$$4OH^-_{aq} \longrightarrow 2H_2O_\ell + 4e^- + O2$$

At cathode:

$$Cr^{+3}_{aq} + 3e^- \longrightarrow Cr_{(s)}$$

Overall reaction:

$$Cr_2(SO_4)_{3(s)} \xrightarrow{water} Cr^{+3}_{aq} \rightarrow 3SO^{-2}_{4(aq)}$$

Q.21 In electroplating of silver, from where Ag + come and where they deposit?

Ans: In electroplating of silver Ag^+ ion come form anode while they deposit at cathode.

Q.22 What is the nature of electrode used in electroplating of chromium?

Ans: In electroplating of chromium, anode is made of antimonial lead while the object to be electroplated acts as cathode.

Exercise Long Question Answers

Q.1 Describe the rules for assigning the oxidation state

Ans: See Q. No. **4** (Subjective Part, Long Questions Answers)

Q.2 Find out the oxidation numbers of the underlined elements in the following compounds.

(a) $Na_2S_2O_4$

(b) $K_2Cr_2O_7$

(c) $AgNO_3$

(d) HNO_3

(e) $KMnO_4$

Ans:

Q.3 How can a non-spontaneous reaction be carried out in an electrolytic cell. Discuss in detail.

Ans: See Q. No. **7** (Subjective Part, Long Questions Answers)

Q.4 Discuss the electrolysis of water.

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.5 Discuss the construction and working of a cell in which electricity is produced.

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.6 How we can prepare NaOH on commercial scale. Discuss its chemistry along with the diagram.

Ans: See Q. No. **12** (Subjective Part, Long Questions Answers)

Q.7 Discuss the redox reaction taking place in the rusting of iron in detail.

Ans: See Q. No. **13** (Subjective Part, Long Questions Answers)

Q.8 Discuss, why galvanizing is considered better than that of tin plating.

Ans: See Q. No. **14** (Subjective Part, Long Questions Answers)

Q.9 What is electroplating? Write down procedure of electroplating.

Ans: See Q. No. **15** (Subjective Part, Long Questions Answers)

Q.10 What is the principle of electroplating? How electroplating of chromium is carried out?

Ans: See Q. No. **16** (Subjective Part, Long Questions Answers)

Unit 8: Chemical Reactivity
Exercise Questions

Exercise Multiple Choice Question Answers:

1. **Metals can form ions carrying charges:**
(a) Uni-positive (b) Di-positive (c) Tri-positive (d) All of them
2. **Which one of the following metals burn with brick red flame when heated in air?**
(a) Sodium (b) Magnesium (c) Iron (d) Calcium
3. **Sodium is extremely reactive metal, but it does not react with:**
(a) Hydrogen (b) Nitrogen (c) Sulphur (d) Phosphorus
4. **Which one the following is the lightest and floats on water:**
(a) Calcium (b) Magnesium (c) Lithium (d) Sodium
5. **Pure alkali metals can be cut simply by knife but iron cannot because of alkali metals have:**
(a) Strong metallic bonding (b) Weak metallic bonding
(c) Non-metallic bonding (d) Moderate metallic bonding
6. **Which of the following is less malleable?**
(a) Sodium (b) Iron (c) Gold (d) Silver
7. **Metals lose their electrons easily because:**
(a) They are electronegative (b) They have electron affinity
(c) They are electropositive (d) Good conductors of heat
8. **Which one of the following is brittle?**
(a) Sodium (b) Aluminium (c) Selenium (d) Magnesium
9. **Which one of the following non-metal is lustrous?**
(a) Sulphur (b) Phosphorus (c) Iodine (d) Carbon
10. **Non-metalsare generally soft, but which one of the following is extremely hard?**
(a) Graphite (b) Phosphorus (c) Iodine (d) Diamond
11. **Which one of the following will not react with dilute HCl?**
(a) Sodium (b) Potassium (c) Calcium (d) Carbon

ANSWR KEY

1	d	3	b	5	b	7	c	9	c	11	d
2	d	4	d	6	b	8	c	10	d		

Exercise Short Question Answers

Q.1 Why reactivity of metals increases down the group?

Ans: Reactivity of metals depends upon its tendency to lose electrons which is dependent upon the size of atoms. Thus reactivity of metals increases down the group because of increasing atomic size and decreasing ionization energy.

Q.2 Why reactivity of metals increases down the group?

Ans: Physical properties of metals:

i. Almost all metals are solids (except mercury).
ii. They have high melting and boiling points.
iii. They possess metallic luster.
iv. They are malleable and ductile.
v. They are good conductors of heat and electricity.
vi. They have high densities.

Q.3 Why nitrogen forms compounds with alkaline earth metals directly?

Ans: Nitrogen forms compounds with alkaline earth metals directly because alkaline earth metals from di-positive cations (M^{++}). They have high charge density and polarization power. They can polarize nitrogen atoms easily and produce stable covalent nitrides with nitrogen.

$$3Mg + N_2 \longrightarrow Mg_3N_2$$

Q.4 Why the second ionization energy of magnesium is higher than the first one?

Ans: Second Ionization energy of magnesium is higher than the first one because after the removal of its electron nuclear charge increases and atomic size decreases. The remaining electrons will be attracted by the nucleus more strongly.

Q.5 How oxygen reacts with group II A metals?

Ans: They are less reactive towards oxygen and they form oxides on heating.

$$2Mg + O_2 \longrightarrow 2MgO$$

Q.6 What is relationship between electro positivity and ionization energy

Ans: Electropositivity depends upon the ionization energy which in turn depends upon size and nuclear charge of the atom. Small sized atoms with high nuclear charge have high ionization energy. Atoms having high ionization energy are less electropositive or metallic.

Q.7 Why electro positivity decreases from left to right in a period?

Ans: Electropositivity decreases across the period from left to right in the periodic table because. Size of atoms decreases due to increase in nuclear charges.

Q.8 How electro positivity depends upon size and nuclear charge of an atom?

Ans: Electropositivity depends upon size and nuclear charge of an atom because when the size of atoms increases, electropositivity increases as it becomes easier to lose electrons. It also depends upon nuclear charge. If nuclear charge increases the electropositivity decreases because it becomes difficult to remove the electrons from outermost shell.

Q.9 Why ionization energies of alkaline earth metals are higher than alkali metals?

Ans: Ionization energies of alkalme earth metals are higher than alkalimetals because the atomic size of alkalme earth metals one smaller and greater nuclear charge.

Q.10 Why silver and gold are least reactive?

Ans: Silver and gold are least reactive because these metals do not lose their electrons easily. They do not have the tendency to make cations.

Q.11 Can pure gold be used for making ornaments? If not why?

Ans: No, pure gold cannot be used for making ornaments because gold is too soft to be used as such. It is always alloyed with copper, silver or some other metal.

Q.12 Why copper is used for making electrical wires?

Ans: Copper is used for making of electrical wires because it is a good conductor of electricity and can easily be drawn into wires.

Q.13 What is the trend of variation in densities of alkali metals?

Ans: Densities of alkali metals increase down the group in the periodic table due to increase atomic mass.

Q.14 Which metal is used for metal work?

Ans: Metal work means objects that are made in an artistic and skilful way. Copper metal is used in metal work because it is easily workable. It is used in many ornaments, plumbing, roafing and other operations.

Q.15 Why magnesium is harder than sodium?

Ans: Magnesium is harder than sodium because in magnesium metallic bonding is stronger, than sodium. Magnesium involves 2 valence electrons in metallic bonding as compared to sodium which involves only one valence electron. Moreover magnesium has smaller atomic size and high ionization energy.

Q.16 Why calcium is more electropositive than to magnesium?

Ans: Calcium is more electropositive than magnesium because calcium has larger size and greater nuclear charge than magnesium and in turn lower ionization energy than magnesium.

Q.17 Why ionization energy of Na is less than Mg?

Ans: Ionization energy of Na is less than Mg because Na has larger size and lower nuclear charge than Mg. It results in less nuclear attraction on valance electron in sodium.

Q.18 Why the ionization energy of Na is more than K?

Ans: The ionization energy of sodium is more than K because down the group electropositive character increases and ionization energy decreases. It becomes easier for potassium to

lose electrons than sodium.

Exercise Long Question Answers

Q.1 Compare and contrast the properties of alkali and alkaline earth metals.

Ans: See Q. No. **5** (Subjective Part, Long Questions Answers)

Q.2 Discuss the inert character of silver and gold.

Ans: Silver and gold are inert metals because they both are very less electropositive and do not lose electrons easily.

Inert Character of Silver:

Silver is a white lustrous metal. Formation of thin layer of oxide or sulphide on its surface makes it relatively un reactive. Under normal conditions of atmosphere, air does not affect silver. It tarnishes in presence of sulphur containing compounds like H_2S.

Inert Character of Gold:

Gold is a yellow soft metal. Gold is very non reactive or inert metal. It is not affected by atmosphere. It is not even affected by any single mineral acid or base. It dissolves only in Aqua Regia. Because of its inertness in atmospheres it is an ornamental metal as well as used in making coins.

Q.3 Why cations are smaller and anions are bigger in size than their respective neutral atoms.

Ans: Small Size of Cation (Positive Ion) than its neutral atom:

Cations are smaller than their corresponding neutral atoms because of two reasons.

i) The removal of one or more electrons from a neutral atom usually, results in the loss of the outer most shell.

ii) The removal of electrons causes an imbalance in proton-electron ration thus a cation has smaller number of electrons than its parent atom with the decrease in number of electrons the magnitude of effective nuclear charge increases, which pulls the electrons cloud of the cation near to the nucleus and thus makes the cation smaller in size than its parent neutral atom.

Examples:

The radius of Na is 186pm whereas ionic radius of cation (Na^+) is 102pm.

Large size of anion (Negative Ion) than its neutral atom

A negative ion is always bigger than its parent atom the reason is that the addition of one or more electrons in the shell of a neutral atom enhance the repulsion between the electron causing the expansion of the shell. The added electrons reduce the attraction of nucleus to the electron that is why the size of anion increases as compared to the neutral atom.

Examples:

Atomic size of Flourine (F) is 71pm whereas anionic size of Fluorine (F^-) is 136 pm.

Q.4 Discuss why hardness and softness of a metal depends upon its metallic bonding.

Ans: The softness and hardness of a metal depends upon the metallic bonding. The strength of the metallic bonds upon the number of valence electrons that each atom contributes for the metallic bonding.

Hardness of a Metal:

Some metals have strong metallic bond due to the greater number of valence electrons in the metal atoms. Such metals are Cord.

Examples:

Magnesium metal has strong metallic bond as compared to sodium metal therefore magnesium is harder than sodium metal.

Softness of a metal

Some metals have weak metallic bond due to the less number valance electrons in the metal atoms. Such metals are soft.

Examples:

Sodium has weak metallic bond as compared to magnesium metal that's why it is soft as compared to magnesium. It has low melting point and can easily be cut with knife.

Q.5 Give the reaction of sodium with; H_20, O_2, Cl_2 and H_2

Ans:

i) Reaction of Sodium with H_2O.

Sodium reacts with water vigorously at room temperature to give strong alkaline solution and hydrogen gas..

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

ii) Reaction f sodium with O_2:

Sodium immediately tarnishes in air giving sodium oxide which forms strong alkali in water.

$$4Na + O_2 \longrightarrow 2Na_2O$$

$$Na_2O + H_2O \longrightarrow 2NaOH$$

iii) Reaction of sodium with Cl_2:

Sodium reacts violently with chlorine at room temperature to give sodium chloride.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

iv) Reaction of sodium with H_2:

Sodium reacts with hydrogen, at high temperature to form sodium hydride.

$$2Na + H_2 \longrightarrow 2NaH$$

Q.6 What are physical properties of calcium metal? Give its uses.

Ans: Physical properties of calcium metal

Following are physical properties of calcium.

1. Calcium is silvery grey and fairy harder.
2. Its density is 1.55g cm^{-3}
3. It is malleable and ductile.
4. It is good conductor of heat and electricity.
5. Its melting point is 851°C and boiling point is 1484°C.
6. Its flame colour is brick red.
7. Its first ionization energy is 590kj mol^{-1} and second ionization energy is 1145kgmol^{-1}.

Uses of calcium

1. It is used to remove sulphur from petroleum products.
2. it is used as reducing agent to produce Cr, u and Zr,

Q.7 Write down the chemical properties of the non-metals?

Ans: See Q. No. **10** (Subjective Part, Long Questions Answers)

Q.8 Compare the physical properties of metals and non-metals

Ans:

Metals	Non Metals
1. All metals are solids except mercury.	1. Non metals are solid, liquid and gases.
2. They have high melting and boiling point.	2. They have low melting and boiling point.
3. They have metallic luster and can be polished.	3. They do not have metallic luster and cannot be polished they have dull surface.
4. They are malleable and ductile.	4. They are not malleable and ductile.
5. They are good conductor of heat and electricity.	5. They poor conductors of heat and electricity.
6. They have high densities.	6. They have low densities.
7. They are usually hard.	7. They are usually soft.

Q.9 How you can compare the softness and hardness of metals?

Ans: Softness and hardness of metals depends upon the strength of metallic bond present in them.

Dependency of Metallic bond

The strength of a metallic bond depends upon the following factors.

i. Charge present on positive metallic ion.

ii. Number of mobile electrons set free by each atom.

Softness of metals:

Metals having weak metallic bond are soft metals, such metals have low melting points,

boiling points, densities etc.

Example:

Sodium is a soft metal due to weak metallic bond. It can be cut with a knife. Its melting point is very low as compared to other metals.

Harness of metals:

Metals having strong metallic bond are hard metals. Such metals have high melting points, boiling points densities etc.

Example:

Magnesium is a hard metal due to strong metallic bond, its melting point is 650°C which is very high as compared to sodium.

Q.10 Give the chemical properties of magnesium and its uses

Ans: Chemical properties of magnesium:

i. Reaction with water

Magnesium reacts with water less rigorously and on heating produces weak base.

$$\mathrm{Mg + H_2O \longrightarrow MgO + H_2}$$

$$\mathrm{MgO + H_2O \longrightarrow Mg(OH)_2}$$

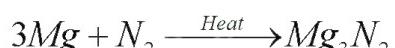
ii. Reaction with oxygen

Magnesium reacts with oxygen on heating and magnesium oxide is formed.

$$2Mg + O_2 \xrightarrow{Heat} 2MgO$$

iii. Reaction with Nitrogen

Magnesium form stable nitride when heat end with nitrogen.

$$3Mg + N_2 \xrightarrow{Heat} Mg_3N_2$$

Uses of Magnesium:

i. Magnesium used in flash light bulbs and in fire works.

ii. It is used in the manufacture of light alloys.

iii. Magnesium ribbon is used in thermite process to ignite aluminium powder.

iv. Magnesium is used as anode for prevention of corrosion. .

Q.11 Write a comprehensive note on the electropositive character of metals?

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Q.12 Compare the ionization energies of alkali and alkaline earth metals.

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)